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COLLIERY MANAGERS' CERTIFICATES*

BY

T. H. BYROM

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LECTURER ON CHEMISTRY, AND ON THE PHYSICS AND CHEMISTRY OF MINING,  
AT THE WIGAN TECHNICAL COLLEGE

**With Illustrations**



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## PREFACE.

THE students of Mining are an ever-increasing multitude, who by present-day requirements are compelled to pay attention to at least the elements of many branches of modern science. The aim of the Author, in these notes relating to the Physics and Chemistry of Mining, is to place before the Mining Student a brief repertory of certain truths and phenomena, such as may be classed under that heading, which it is hoped may materially assist him in his studies.

As a general rule, text-books on Physics and Chemistry—though often excellent for the purposes for which they are written—deal largely with subjects which are neither interesting nor useful to the student of Mining, and either treat very slightly or omit altogether those points which are of special interest and importance to him. Further, there are very few books in which the principles of those sciences are dealt with in one volume.

The composition of the air; the importance of, and reason for, an abundant air-supply; the constitution and properties of the gases to be met with in the mine, some of them most deadly either as regards explosive properties or poisonous effects; the behaviour of gases under varying conditions of temperature and pressure; the indications of the barometer; the humidity of the air of the mine, and the dangers of coal dust: these (with many other) matters ought to have the most careful attention in a student's training and preparation for responsible work in the mine; and the Author trusts that the treatment of

such matters which is here given may be found of substantial practical utility as an elementary first course.

A series of questions, many of them taken from papers set at the Colliery Managers' Examinations and at the Board of Education's Examinations in Mining, has been appended to the volume.

The Author desires to acknowledge here, with many thanks, his indebtedness to Mr W. J. Orsman for information furnished in connection with the chapter dealing with Explosives; to Mr G. H. Winstanley for help and advice in several matters dealing more especially with practical mining; and to Messrs Philip Harris & Co., Birmingham, for the loan of many of the blocks used for the illustrations.

T. H. B.

WIGAN,  
*May 1905.*

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## CHAPTER I.

### THE ATMOSPHERE.

WE are surrounded on all sides by an invisible gaseous mixture, commonly known as the *atmosphere* or *air*; and a study of the properties, physical and chemical, of this mixture, which obeys the laws relating to gases in general, will be interesting and instructive. Every so-called empty vessel is really full of air, and we are reminded of its presence by many things in our daily life; for instance, the resistance one experiences in walking against a breeze, or the disastrous effects of a wind storm, due to the air having been in a state of violent disturbance. From the very earliest periods, it must have been evident to man that the invisible "something" was necessary to life, as when deprived of it life became extinct, and the important and far-reaching effects of this atmosphere of ours are evident on every hand. The various changes of weather are due to it; soils are formed by the disintegration or crumbling of rocks by means of it; all the phenomena of sound would be impossible without its aid; by its behaviour the sun's rays, instead of scorching up the face of the earth, are modified in passing through the stratum or layer of air surrounding the globe.

**The Air possesses Weight.**—This may be proved by weighing a circular globe full of air (Fig. 1) and then attaching to it an air-pump, whereby the air may be removed, and afterwards re-weighing. There will be found a considerable decrease in weight.

A cubic foot of air, at a temperature of 0° Centigrade

and a pressure of 30 inches of the barometer, weighs .0807 lbs.; that is to say, about  $12\frac{1}{2}$  cubic feet weigh one pound, and consequently the weight of air in a room 16 ft. by 14

ft. by 10 ft. will be, say, 180 lbs. In consequence of this weight, and the fact that the atmosphere extends to a height of several miles, it is to be expected that a considerable amount of pressure will be exerted by it upon all objects on the earth's surface.



FIG. 1.—Apparatus for weighing air.

### **Atmospheric Pressure.**—This amounts to about

14.7 lbs. on every square inch of surface. Even on our bodies this pressure is always present, and totals up to several tons pressure, but being equal on all sides and from all directions, it does not cause inconvenience. If, however, the air from the *inside* of a vessel be removed by some means, such as an air-pump, unless the vessel is very strongly made and (preferably) circular in shape, the whole substance of it will be crushed in by the atmospheric pressure.

A striking experiment in proof of this may be made by taking a square tin canister, soldering a small tube in the lid, then soldering the lid securely on to the box. Introduce a little water into the canister and place it over a Bunsen gas burner, or the flame of a spirit lamp, until steam begins to issue from the tube. When this takes place remove the source of heat and quickly cork up the tube tightly. As the tin and its contents (steam) cool down, the steam will become condensed and a partial vacuum be formed inside; the resistance inside to the atmospheric pressure outside will get less and less, until finally the tin is crumpled up almost flat.

**The Barometer.**—That the atmosphere exerts pressure is also shown by means of the barometer, an instrument the construction and principle of which may be readily shown as follows: Take a glass tube 32 or 33

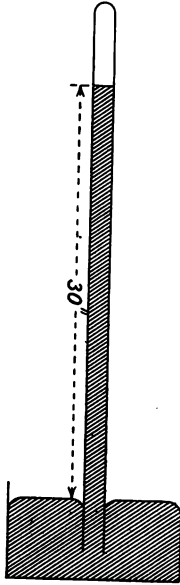


FIG. 2.—Diagram illustrating construction of barometer.



FIG. 3.—Common form of pit barometer, with thermometer attached.

inches long, and sealed at one end. By means of a small funnel, fill the tube with the fluid metal mercury, commonly called quicksilver, and placing the thumb on one end, turn the tube upside down, and bring the end under the surface of mercury contained in a small trough (Fig. 2).

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Remove the thumb when the end of the tube is well under the surface of the mercury in the trough. The mercury, which before completely filled the tube, now falls until its level is about 2 or 3 inches from the sealed end, and there it remains stationary. This space above the mercury is a vacuum (except for a little vapour of mercury). On measuring the distance between the level of the mercury in the trough and that in the tube, it will be found to be about 30 inches (corresponding to the height of the barometer at the time of the experiment). The column of mercury is supported entirely by the atmospheric pressure on the surface of the metal in the trough (Fig. 3).

If the bore of the tube had an area in section equal to 1 square inch, there would be 30 cubic inches of mercury in a column 30 inches in height. Now 1 cubic inch of mercury weighs practically  $\frac{1}{2}$  lb. (.491), therefore the whole column would weigh  $30 \times \frac{1}{2} = 15$  lbs. But this column balances, or is balanced by, a column of air of the same area. Hence the weight or pressure of the air, on an area of 1 square inch, is equal to 15 lbs.

Mercury is the only metal which is liquid at ordinary temperatures, and on account of its weight it is eminently suitable for use in barometers. Water might be used for a rough kind of barometer, but there would be many objections, and a column of  $30 \times 13\frac{1}{2}$  inches in height would be necessary ( $13\frac{1}{2}$  being the density of mercury compared with water). Glycerine is the liquid used in some barometers. Its specific gravity is 1.26, and a column of about 27 feet in height is necessary. It possesses the great advantage that very slight variations in the atmospheric pressure are more readily indicated, being magnified practically ten times, as compared with an ordinary mercury barometer.

Another form of barometer is also in common use, named the aneroid (Fig. 4). In principle it consists of a flattened cylinder of corrugated metal, from which the air has been taken by means of an air pump. Fluctuations in

the pressure of the air cause the cylinder to be flattened slightly, or to expand slightly, and these movements are imparted to suitable mechanism and indicated by a pointer on the face of the instrument.

**Importance of the Barometer to the Mine Manager.**—The barometer is an instrument of the greatest importance to the mine manager. By its means he is able to detect variations in the pressure of the air which will affect the ventilation of the mine. The greater the atmospheric pressure—that is, the higher the barometer column—the less danger of fire-damp in a coal mine, a *lessened* pressure denoting *more* likelihood of escaping gas from the fissures and cracks in the coal strata and from the coal itself.

The pressure of air on a mountain top is less than it is at the sea-level, and the higher one ascends, the less the pressure becomes. It is found that a difference of 900 feet in height means a rise or fall of 1 inch of the mercury in the barometer. Consequently the pressure in the workings of the mine will be *greater* than at the surface, to an amount depending upon the depth of the mine.



FIG. 4.—Aneroid barometer.

**The Siphon.**—If a U-shaped tube, with one limb longer than the other, be taken and filled with water, no water will run out so long as the tube is sealed by the finger at one end, and the tube may be safely turned upside down without losing any of the liquid. The atmospheric pressure prevents it from coming out. Now if the open

and shorter limb be placed in water and the sealed end be placed at a lower level than the surface of the water, on removing the finger the liquid will flow out of the longer limb, and continue to do so until the vessel of water is either emptied, or its surface comes below the short end of the U tube. Such an arrangement is called a siphon, and it depends for its action on atmospheric pressure.

The column of liquid in the long limb weighs more than that in the short limb, and consequently runs out at the end of the long limb; the liquid in the shorter limb having to follow it, or else a vacuum would be formed. The effect is thus due to atmospheric pressure forcing up liquid in the short limb to compensate for that running out of the long limb.

The siphon is a very useful appliance in a laboratory for various purposes where such action of the atmosphere is to be utilised, as in separating a clear liquid from a sediment which has settled in it. Its practical application in the mine has limitations. It may be used for raising water over an elevation and delivering it at some lower level, or for draining water from a higher level to the pump, but it can never be used to transfer water from a lower to a higher level. Since its action depends upon atmospheric pressure, the vertical height over which water can be raised is about 34 feet, *theoretically*. In actual practice it is less than this, and rarely can much more than 20 feet be effectually dealt with, on account of leakage of the pipes and loss through friction.

The siphon on such a scale as this is constructed of a series of pipes joined together, and slight leakages at the joints is almost unavoidable. To start the siphon, the ends of the pipes are plugged and the pipes filled with water through an opening at the top of the bend. This is then sealed as tightly as possible, and the plugs or stoppers removed from the ends of the pipes, when a similar action goes on to that which takes place when using a glass tube

with unequal limbs. The end of the shorter limb is closed by a ball valve when the water gets below its level, thus keeping the pipes full of water.

The siphon is not often used in mining operations, but it may be useful sometimes where pumping machinery is not available.

The action of an ordinary suction pump depends also on atmospheric pressure. Before the pump is put in action the barrel and pipe contain air only, and the level of the water inside and out is the same. On moving the handle, a partial vacuum is caused under the piston, and the pressure of the atmosphere forces water up into the working barrel.

## CHAPTER II.

### LAWS RELATING TO THE BEHAVIOUR OF GASES.

WE have now to consider the laws relating to the behaviour of gases when subjected to certain conditions.

**Boyle's Law.**—When any gas, or mixture of gases, is subjected to pressure, its bulk or volume is affected. If the pressure be *increased*, the volume is *decreased*; and *vice versa*, if the pressure be *decreased* then the volume is *increased*. Expressed more concisely, we should say that “the volume of gases varies *inversely* as the pressure to which they are subjected.” This is known as Boyle's law, and it assumes the *temperature* to be constant.

It may be clearly proved by means of a bent glass tube sealed at one end, or two pieces of tube connected by rubber tubing, as shown in Fig. 5. The limb is closed by pouring a little mercury down the mouth of the open limb *a*. The column of air in the short limb *b* is now at the same pressure as the outside air is. More mercury is now poured in the longer limb until a height of 30 inches is attained, and the difference in volume of gas in the short limb is noted. It will be found to have decreased to *half* its

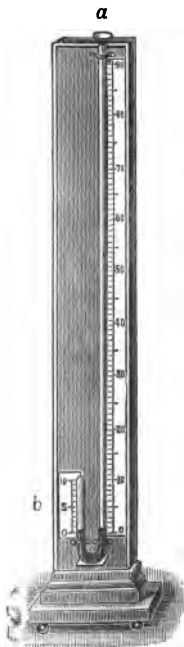


FIG. 5.—Boyle's tube.

original volume. That is to say, by adding another atmosphere's pressure (which we know to be equal to 30 inches of mercury approximately) the volume of air has become halved. We have therefore *doubled* the pressure and *halved* the volume.

**Effect of Temperature on Volume of Gases.—Charles's Law.**—The temperature is an important matter in dealing with the volume of gases, and the physicist Charles, who investigated the subject, found that gases have a definite and regular increase in volume for a definite increase in temperature, *the pressure remaining the same*. He proved that a gas increases in volume  $\frac{1}{273}$  of its bulk for every 1° C. increase in temperature.

Plainly expressed, we should find, therefore, that 273 cubic feet of air, when raised in temperature 1° C., would become 274 cubic feet.

$$\left( \frac{1}{273} \text{ of } 273 \right) + 273 = 274.$$

The opposite also applies, namely, for a decrease in temperature a corresponding decrease in volume occurs in the above proportion.

Suppose that a certain volume of air measures 50,000 cubic feet when the barometer indicates a pressure of 30 inches. What will be the volume if the pressure falls to 29 inches? If the temperature be also increased, say from 15° C. to 25° C., how will the volume be changed, and to what extent?

Consider the pressure first. This becomes less; therefore the volume becomes greater, and in this proportion

$$50,000 \times \frac{30}{29} = 51,724 \text{ cubic feet.}$$

Next consider the temperature, which has increased, and therefore causes increased volume.

From Charles's law this will be seen to be

$$\frac{51,724 \times (273 + 10)}{273} = 53,618 \text{ cubic feet,}$$

or (combining the two calculations or corrections for pressure and temperature in one)

$$\frac{50,000 \times 30 \times 283}{29 \times 273} = 53,618 \text{ cubic feet.}$$

That gases expand on heating may be clearly proved by the following experiment: Procure a small glass flask, and fit it with a good cork, through which a clean cut hole has been bored to receive a piece of glass tubing bent twice at

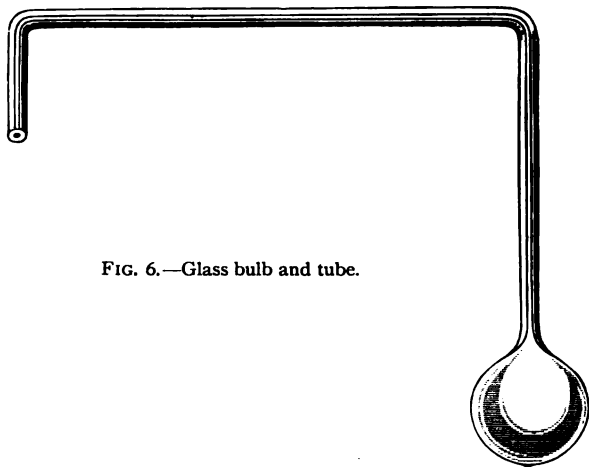


FIG. 6.—Glass bulb and tube.

right angles. The end of this tube dips under water in a glass vessel. On applying heat to the flask, even so slight as the warmth of the hand, the air will expand and bubble through the water in the vessel. More heat causes greater expansion, and on allowing the flask to cool down again the water will be drawn back into the tube and probably into the flask as well, showing that decrease of heat causes decrease of volume. Instead of a flask, the bulb and tube, as shown in Fig. 6, can be used.

Gases as a rule are so light, compared with solids and liquids, that we are accustomed to regard them as having no weight. As we have seen, the air itself possesses

weight; and it may be proved by using the same apparatus that the other gases with which we are acquainted have each of them densities and specific gravities peculiar to themselves. When the density of a gas is spoken of, it is understood to be compared with hydrogen gas as a standard, taken as 1. Thus the density of air is 14.47, and of oxygen 16.0. That is, air and oxygen are respectively 14.47 and 16 times as heavy as hydrogen.

When the specific gravity of a gas is mentioned, it is understood that the comparison is made with air as a standard. Thus the specific gravity of  $\text{CO}_2$  (black damp) is 1.5, and of  $\text{CH}_4$  (fire damp) .56, one being  $1\frac{1}{2}$  times as heavy, the other approximately half as heavy, as air.

The densities or specific gravities of all gases are affected by the temperature. If the temperature be increased the densities are lowered, and if the temperature be decreased the densities are increased. The pressure also affects the volume, and therefore the weight, if the gas be free to expand or contract. Therefore the comparison of all densities and specific gravities are understood to have been made at one standard temperature and pressure, namely—

0° Centigrade and 760 millimetres barometer.

The density or specific gravity of gases may also be determined in the apparatus shown in Fig. 7. These U tubes are made to contain a certain volume of gas, usually 25 c.c., but their volume should always be checked by filling with mercury and measuring the volume of mercury used. The gases whose density is to be determined are first very carefully *dried* and freed from impurities, and the purified gas allowed to stream through them until all air or other gas has been expelled, then carefully weighed—the

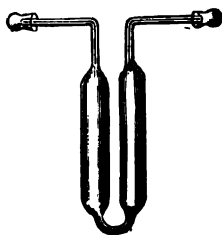


FIG. 7.—Tube for determining densities of gases.

temperature and pressure being noted at the time (avoid handling too much with warm hand; carry carefully by the capillary tubes)—and afterwards corrected to the standard temperature and pressure. In this way comparative weights of equal volumes of different gases at the same tempera-

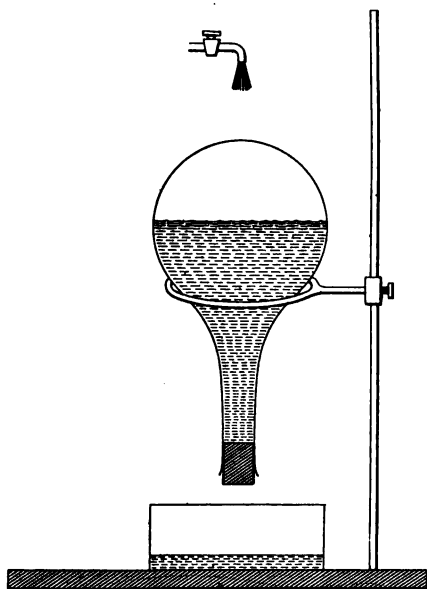


FIG. 8.—Apparatus to show lowered boiling point.

ture and pressure are obtained, from which their density may be got.

**The Effect of Pressure on the Boiling Point of Water and Liquids generally.**—The effect of pressure upon the boiling point of a liquid is a very remarkable and important one. The result of observation goes to prove that the boiling point of liquids is raised by increasing the pressure on the surface of the liquid, and lowered by decreasing the pressure. The boiling point is always

arrived at when the *tension* or *elastic force* of the vapour (say steam) is able to overcome the atmospheric pressure, or other pressure to which the liquid may be subject. It follows from this that the temperature of the water in a steam boiler is much higher than the temperature of water boiling at ordinary atmospheric pressure (*i.e.* 100° C. or 212° F.); and conversely, if the pressure be lessened, then the boiling point is lowered.

This may be made the subject of an interesting experiment, as follows:—

Take a round-bottomed flask (Fig. 8) (so that it will stand a fair amount of pressure) and half fill it with water, placing over the flame to boil. Have fitted to it a sound cork, through which passes a tube having a pinch-cock attached. When the water boils and steam issues from the tube, close the pinch-cock and remove the source of heat. Turn the flask upside down and place on a suitable support. There will be now a space above the water filled with steam, and if a little cold water be sprinkled on the flask, condensation takes place, a partial vacuum is produced, and consequently diminished pressure on the surface of the water, and the latter commences to boil briskly, although its temperature is by this time much below 212° F. or 100° C. Also, by reducing the pressure on the surface of water by means of an air-pump, water may be made to boil at a much lower temperature than under ordinary conditions.

At high altitudes, where the atmospheric pressure is less than at sea level, the boiling point is correspondingly lower, and special apparatus has to be used for cooking purposes, so that the chemical changes necessary in cooking the food material in use on the occasion may be effected. At a height of  $3\frac{1}{2}$  miles the density of the air is only one-half what it is at the sea level, and at 7 miles only one-third.

The presence of solids dissolved in the water raises the boiling point. Sea water, for instance, does not boil at

212°, and it is often an advantage to add certain salts to water in order to maintain a uniform temperature, higher than the normal boiling point (212°).

**Heat is a Form of Energy.**—It has long been known that heat could generate mechanical power, and also that mechanical power could generate heat. Dr Joule actually measured the amount of heat that was given to a certain weight of water by a certain amount of mechanical power, and from this the amount of work which could be got from a certain amount of heat; and he proved that the quantity of heat produced is always equal to the quantity of work expended in producing it.

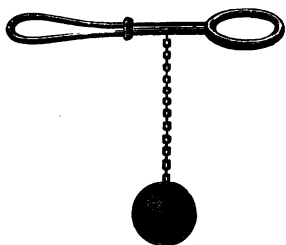


FIG. 9.—Apparatus to test expansion of metal ball.

The following changes usually accompany the application of heat to matter. The solid expands, then changes to liquid; the volume of the liquid expands, the liquid becomes converted into gas, and the gas or vapour expands. That heat causes cubical expansion may be shown

by means of a metal ball, which just passes through a ring when at ordinary temperatures (Fig. 9). On heating the ball, it is found that it will no longer pass through the ring because of increased volume, due to heat.

Each metal has its own *co-efficient of expansion*—that is to say, it expands a definite amount for a definite increase in temperature. The co-efficient of expansion of iron, for instance, is  $\cdot0000122$ , which means that for every 1° C. heated, it increases *in length* (linear expansion)  $\cdot0000122$  of its original length.

For example.—A bar of iron 20 feet long is heated from 15° C. to 95° C. How much will it increase in length?

For the 80° C. increase in temperature it will expand  $\cdot0000122 \times 80 = \cdot000976$  of its length.

20 feet = 240 inches.

$\therefore \frac{976}{1,000,000} \times 240 = \cdot 234$  inches increase in length, or nearly  $\frac{1}{4}$  of an inch.

The total increase in volume or bulk (cubical expansion) will be three times this—namely,  $\cdot 702$  cubic inches.

Fig. 10 is an illustration of the apparatus used to test the linear expansion of a metal bar when heated.

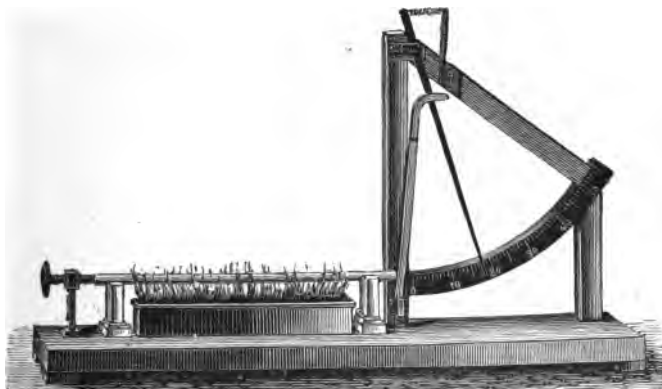


FIG. 10.—Apparatus to test linear expansion of metal rod.

**The Effects of Heat on Solids** are put to many practical applications. Iron rods are made to brace up buildings, and other structures, by being first heated and then screwed up tightly at the ends, and allowed to cool and contract or shrink. This force of contraction is irresistible, as may be shown by using the apparatus shown in Fig. 11, in which a heated iron bar is placed and screwed tightly up; on cooling, the bar is broken in pieces by the contractile force.

Iron tyres are shrunk on to wheels by being heated first and allowed to cool on the wooden rim, the spokes, rim, and hub being thus firmly bound together. In the metal work of bridges, and other structures, allowance

must be made for expansion of material due to heating (see problem on expansion); and the sections of rails on a railway are always laid a small distance apart at their ends: otherwise the expansion on a hot summer's day would be sufficient to cause a disaster, owing to twisting of the track.

The reason why thick glass vessels will not stand heat is, therefore, easily explained; the glass on heating becomes expanded, but being a bad conductor (see "Conduction") is not able to transmit the heat quickly enough, and the two portions of the glass are in unequal strain and consequently crack, since the glass is not elastic enough

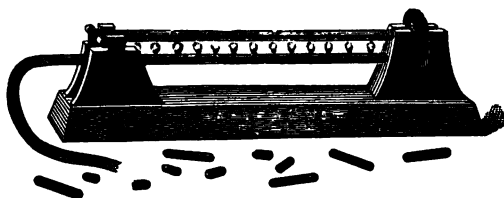


FIG. 11.—Apparatus to show effect of contraction on heated metal bar.

to prevent this. *Thin* glass vessels, such as are used in chemical operations, cause no such trouble, and liquids are constantly *boiled* in them.

Heat is conveyed from one substance to another in several ways—by conduction, by convection, and by radiation.

By **Conduction** the heat is transferred from one particle of the substance to another, the particles having a higher temperature giving up heat to those near to them having a lower temperature, until the heat is equalised.

Substances vary much in their power of conducting heat. Good conductors are those which readily transmit heat from one part of their substance to another. This may be shown by taking bars of iron and copper, for instance, of the same dimensions, and fixing on them at

intervals, by means of paraffin wax, small wooden balls (Fig. 12). On heating the bars, as shown in the figure, the metals take up the heat and transfer it according to their conductivity, the wax melts, and allows the balls to drop as the temperature rises. It will be found that the balls will drop off the copper bar before the iron, thus proving its greater conducting power for heat.

Gases and liquids are bad conductors of heat, as are many solids, except the metals. Asbestos is a particularly bad conductor, and this, in conjunction with its power of resisting fire, makes it a valuable material for fire prevention and fire-proofing.

In **Convection** the heat is carried from one portion to

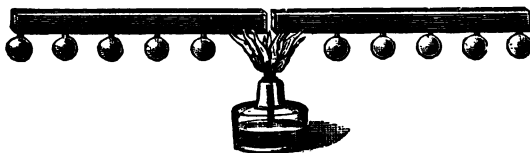


FIG. 12.—Experiment to show conducting power of metals.

another by *actual movement* of the particles. This may be well shown by sprinkling sawdust in a flask partly filled with water, and then applying heat, whereupon the particles of sawdust move about and show the direction of the convection currents. Convection can only take place in liquids and gases.

**Radiation.**—When we stand near a source of heat, say a fire or furnace, we feel the warmth which is *radiating* from that source. If a screen be placed between, the difference is at once noticed, which proves that the sensation is not due simply to the air itself being hot. Radiation is going on in a greater or lesser degree from the surfaces of all bodies, and the intensity varies according to the intensity of the source, and also with the distance from it, and the radiating power of the heated surface or body.

**Heat can be reflected**, just as light can. This may be proved by means of two concave mirrors arranged as shown in Fig. 13. In the focus of one of the mirrors is placed a source of heat, such as a hot metal ball; in the focus of the opposite mirror is placed a small piece of phosphorus or some inflammable material, the latter (it will be seen) being readily set on fire by the reflected and focussed heat. The heating effect produced by the sun's rays through an ordinary double convex lens is well known.

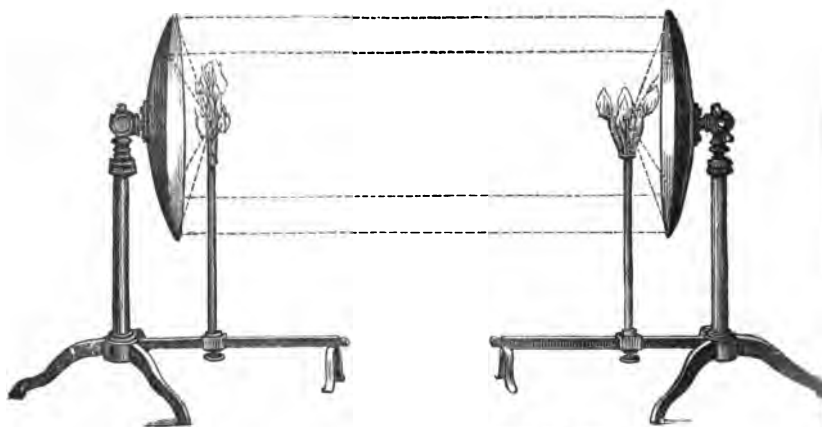


FIG. 13.—Arrangement of mirrors, so as to focus heat.

The reflection from plane (flat) surfaces may be shown by means of an air thermometer (see description, p. 22). Bodies which best reflect heat naturally *absorb it least*, and substances which take up heat most readily are the worst reflectors; consequently light-coloured clothing in summer is cooler than black. On the other hand, it has been proved that bodies which best absorb heat also *radiate* it best, and on this account vessels which are required to hold or retain their heat should be kept brightly polished, so as to prevent radiation as much as possible.

**Heat also promotes chemical action**, and affects

the conductivity of metals for heat and electricity. Bodies vary greatly in the *amount* of heat they can take up—that is, in their *capacity for heat*, generally spoken of as their *specific heat*.

Water is taken as the standard whereby specific heats are compared, since it has the highest capacity for heat of all known substances; and the standard being 1, the specific heat of mercury (quicksilver), for instance, is  $\cdot 033$ , which is practically  $\frac{1}{30}$ . The meaning of this is that it requires thirty times as much heat to raise one pound of water  $1^\circ$  as it does to raise one pound of mercury  $1^\circ$ . This explains the great efficiency of hot water as a means of heating the air in a room.

Thus: The specific heat of air is  $0\cdot 237$ , that of water being 1. Hence 1 lb. of water in cooling through  $1^\circ$  C. heats  $\frac{1}{0\cdot 237}$  lbs. =  $4\cdot 2$  lbs. of air  $1^\circ$  C.

Again: Water is 770 times heavier than air; therefore 1 cubic foot of water in cooling through  $1^\circ$  C. heats  $(770 \times 4\cdot 2) = 3234$  cubic feet of air  $1^\circ$  C.

**Latent Heat.**—If a pound of water at  $80^\circ$  C. be mixed with a pound of water at  $0^\circ$  C., the result would be two pounds of water at  $40^\circ$  C.; but if one pound of water at  $80^\circ$  be mixed with one pound of *ice* at  $0^\circ$ , the result would be two pounds of water at  $0^\circ$ . Neither the temperature nor the volume of the ice has been increased; it has simply become liquefied to water, and the  $80^\circ$  heat has become *latent* or has apparently disappeared. Thus  $80^\circ$  is known as the latent heat of water. It is the amount of heat necessary to convert ice at  $0^\circ$  into water at  $0^\circ$ .

Again, when heat is applied to water the temperature rises until the boiling point is reached, when the continued application of heat causes *no increase in temperature*. A thermometer placed in the vapour (steam) does not indicate more than  $100^\circ$  C., however much heat is applied to the water. The heat being thus expended is being used

up in converting the *water* at 100° C. into steam at 100° C., and this is known as the *latent heat* of steam. If the steam be passed into water, its latent heat is converted into *sensible heat* again.

For example: 20 lbs. of steam at 100° C. are passed into 200 lbs. of water at 10° C., and condensed. The temperature of the total water finally is 67° C. What is the latent heat of the steam?

*Heat lost by steam.*                                 =     *Heat gained by water.*

If  $l$  represents the latent heat, then—

$$\begin{array}{rcl}
 20\ l + 20(100 - 67) & & 200(67 - 10) \\
 20\ l + 660 & = & 11400 \\
 \therefore 20\ l & = & 11400 - 660 \\
 \therefore l & = & 537
 \end{array}$$

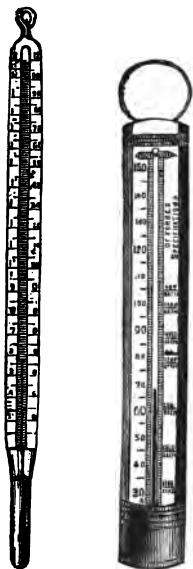
Hence by this experiment the latent heat of steam has been found to be 537.

**Heat and its Measurement.**—Instruments used to measure the *intensity* or *degree* of heat are called thermometers, and those used to measure the *amount* or *quantity* of heat are called calorimeters.

The ordinary thermometer (Figs. 14 and 15) consists of a glass stem through which passes a capillary (*i.e.* very fine) tube, and has a bulb at one end which serves as a reservoir for the liquid employed, mercury or alcohol, which latter remains liquid at a much lower temperature than mercury does. Mercury is used because it is a liquid, and remains so for a fair range of temperature; it is easily heated and expands regularly. This bulb and tube are filled with mercury by alternately heating and cooling the bulb and keeping the end of the fine tube under mercury. The tube is sealed up by a blowpipe flame when a suitable length of it is filled.

It is next necessary to graduate the stem, which is done by taking two fixed points, the freezing and boiling points of water. By placing the thermometer in a suitable

apparatus, these points are obtained and their position marked on the stem. Fig. 16 shows the apparatus used to determine and mark the point at which water boils. The thermometer bulb is surrounded by steam from the water which is kept boiling in the vessel. The freezing point is determined by surrounding the bulb by melting



FIGS. 14, 15. — Common forms of thermometers.

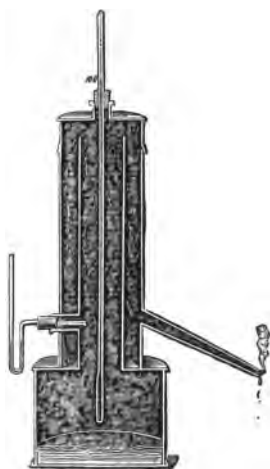


FIG. 16. — Apparatus for graduating thermometers at the boiling point of water.

ice. The space between the two is divided according to the scale decided upon. If the Centigrade graduation, the freezing point is marked  $0^{\circ}$  and the boiling point  $100^{\circ}$ , and the stem is divided into 100 equal parts between these two points. If the Fahrenheit scale, then the freezing point is made  $32^{\circ}$  and the boiling point  $212^{\circ}$ , and the intervening space divided into 180 equal parts. The other system of graduation due to Reaumur is never used in this country. On it the freezing point is  $0^{\circ}$  and the boiling point  $80^{\circ}$ .

A delicate and useful means of measuring heat effects is that devised by Sir John Leslie. It depends for its action upon the expansion or contraction of a volume of air. The instrument, as shown in the illustration (Fig. 17), consists of two bulbs connected by a glass tube, and fitted on a stand. The bulbs contain air, and the glass tube is partly filled with coloured liquid. On bringing a source of heat near

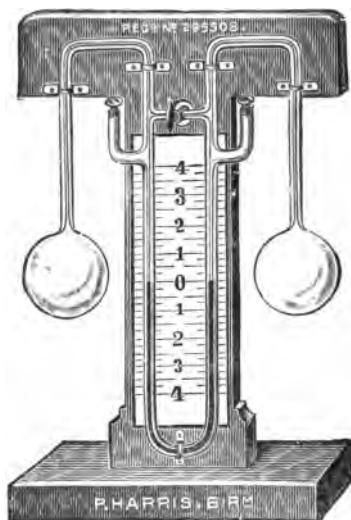


FIG. 17.—Leslie's air thermometer.

one of the bulbs, or by placing the warm hand on it, expansion of the air takes place and causes the liquid column to move. The stem is graduated into divisions according to experiments made at known temperatures.

It is often desirable to know the lowest temperature, and also the highest, that has prevailed during a certain time, say twenty-four hours. An instrument called the maximum (greatest) and minimum (least) thermometer has been constructed for this purpose. As shown in the illustration (Fig. 18), it consists of two thermometers of

special shape fitted on a suitable support. In one the liquid is mercury, in the other alcohol, the former being the maximum thermometer, and the latter the minimum. The maximum has a small piece of iron rod moving freely in the tube; as the mercury expands it pushes this small rod before it, and it remains at the furthest or highest point thus reached, and therefore shows the maximum temperature reached in a given period. The minimum thermometer has a small glass tube moving freely in it. When the alcohol expands, it simply *flows through* this index tube without moving it; but when it contracts the small index is drawn back by the receding liquid, when

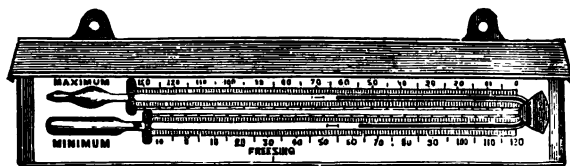


FIG. 18.—Maximum and minimum thermometers.

the end of the alcohol column is reached, by force of cohesion or capillary action. In this way the index always remains at the lowest temperature touched by the alcohol thermometer.

It has been pointed out (p. 9) that according to Charles's law 273 volumes of gas (or air, of course), when raised in temperature  $1^{\circ}$  C. become 274 volumes—that is,

$$273 + \left( \frac{1}{273} \text{ of } 273 \right).$$

Assume that a gas occupies 273 volumes, measured at  $0^{\circ}$  C.

This would become 274        „        „        at  $1^{\circ}$  C.

275        „        „        at  $2^{\circ}$  C.

On the contrary 273 volumes measured at  $0^{\circ}$  C.

would become 272        „        „         $-1^{\circ}$

271        „        „         $-2^{\circ}$

171        „        „         $-100^{\circ}$

and theoretically, nil        „        „         $-273^{\circ}$

So that, in theory, the gas would disappear altogether at a temperature of  $-273^{\circ}\text{C}$ . This is impossible, practically. As a matter of fact, long before this temperature is reached, the gas becomes either liquefied or solidified, and Charles's law fails to operate. Nevertheless, the point  $-273^{\circ}\text{C}$ . is always spoken of as the *absolute zero* of temperature.

**Conversion of Degrees.**—For the conversion of degrees Fahrenheit into degrees Centigrade, and *vice versa*, the subjoined rules may be followed:—

To convert  $\text{F}^{\circ}$  into  $\text{C}^{\circ}$  subtract 32 from the original number, and divide the remainder by 1.8. Thus  $176^{\circ}\text{F} - 32 \div 1.8 = 80^{\circ}\text{C}$ .

To convert  $\text{C}^{\circ}$  into  $\text{F}^{\circ}$  multiply by 1.8 and add 32. Thus  $80^{\circ}\text{C} \times 1.8 + 32 = 176^{\circ}\text{F}$ .

Taking the melting point of copper as  $1100^{\circ}\text{C}$ . what is this on Fahrenheit scale?

$$\frac{1100 \times 9}{5} + 32 = 2012^{\circ}\text{F}.$$

The ignition point of an explosive mixture of fire-damp and air is  $1230^{\circ}\text{F}$ . What degree does this represent on the Centigrade scale?

$$\frac{1230 - 32}{1.8} = \frac{1198}{1.8} = 665^{\circ}\text{C}.$$

(It should be noted that  $\frac{9}{5}$  is equal to 1.8.)

## CHAPTER III.

### THE DIFFUSION OF GASES.

THE property that all gases have, of diffusing or spreading one into the other, is a very important one in Nature, and especially so to those connected with mining.

**Diffusion of Gases, how illustrated.**—It is well illustrated by filling two gas jars—one with carbon-di-oxide, the other with hydrogen; one a very heavy gas, the other extremely light. Then place the jars mouth to mouth, the one containing the hydrogen being uppermost. It would be expected that, owing to the great difference in their densities (carbon-di-oxide is twenty-two times heavier than hydrogen), the two gases would remain as placed, the light one resting on top of the heavy one, but this is not the case. It will be found in a very short time that some of the hydrogen has diffused into the carbon-di-oxide, and also that the latter has diffused into the hydrogen; and before long the gases become so intimately mixed, that the composition of the gas in the lower jar is the same as that in the higher jar. Further, these gases will never separate out again into a heavy and a light layer, as they were at the beginning of the experiment.

This effect takes place in the case of all gases or mixtures of gases, whatever their difference in density; and strangely enough, the greater the difference between the densities of the gases, the more rapidly the diffusion goes on.

A knowledge of this law explains why it is that a heavy gas like carbon-di-oxide is not always found on the floor of a mine, but equally diffused throughout the atmosphere. Fire-damp also, although collecting at first in the highest

parts of the roof, soon diffuses out into the heavier atmosphere, and forms an explosive mixture unless sufficiently diluted with air.

**Testing Diffusion of Gases.**—The principle of gaseous diffusion is well illustrated by the apparatus shown in Fig. 19.

It consists of a porous cell (of the kind used in electric batteries), fastened to one limb of a U-shaped tube. The tube is filled with water to the same level in both limbs. On bringing a glass vessel containing hydrogen, or coal-gas, or fire-damp, over the porous cell, the gas in the vessel diffuses through into the porous cell more quickly than the air can escape, and consequently produces pressure inside the cell, which is shown by the forcing out of the water from the open end of the U tube.

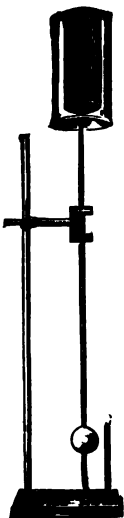


FIG. 19.—Apparatus for diffusion experiment.

An ingenious application of the principle has been made in the case of Ansell's fire-damp indicator. On placing the apparatus, consisting of porous cell and U tube, in a suspected atmosphere, if fire-damp be present it diffuses through more rapidly than the air can escape, and causes pressure, which is made to act on a column of mercury (instead of water). The mercury rising in the open limb makes electrical contact with a platinum wire fused through the side of the tube, and completes an electric circuit, in which is a bell. The other terminal is a wire in the bend of the U tube.

**The Hygrometer and its Use.**—The importance of a certain amount of moisture in the air of a mine cannot be over-estimated, since the presence of coal dust in the atmosphere seriously increases the likelihood of an explosion.

Mr Henry Hall's experiments with coal dust proved that

when the flame from a blown-out gunpowder shot, or the ignition of a small quantity of fire-damp, passes through an atmosphere containing even a moderate quantity of dry coal dust, the dusty atmosphere will explode with great violence, and that the explosion passes on throughout any length of such atmosphere, increasing in force as it progresses. Further, that coal dust from several seams in various districts is almost as sensitive to explosion as gunpowder itself.

The probable cause of this explosive action, chemically, is the extremely fine condition of the coal dust, rich in carbon and hydrogen (see later for their properties), only waiting opportunity to combine with oxygen from the air. Such opportunity is afforded by an increase in temperature from any cause such as a blown-out shot or slight ignition of fire-damp, more especially if carbonic oxide be present also.

The mine manager, therefore, should have at his disposal an instrument for indicating the amount of watery vapour, humidity, or dampness in the air. The presence of aqueous vapour in the air is not always readily recognised by the senses, but can be clearly proved to be present by its action on various chemical substances which greedily absorb it. In summer, though the air may *feel* dry, there is really more moisture in it than on a winter's day, when air is said to be damper, the warmth of the atmosphere helping it to hold more vapour. Sulphuric acid (oil of vitriol) on exposure to the air soon increases in bulk; and the solid, calcium chloride, takes up water so readily that it liquefies itself in a very short time. Another useful agent is known as anhydrous copper sulphate. When dry it is white, but on absorbing moisture becomes blue, so that it readily shows the presence of moisture in air or any gas. By using suitable substances of this kind, and noting the increase in weight when a known volume of air is passed over them, the amount of moisture in a given volume of air is readily estimated.

By means of the apparatus shown in Fig. 20, the amount of moisture in the atmosphere may be measured. The tubes *a* contain calcium chloride, the volume of the aspirator *b* is known, and the air is filtered through cotton wool in the tube *c* to separate dust. The gain in weight of the tubes *a* represents the moisture in the volume of air drawn through them by the aspirator.

This method, however, would be too lengthy and inconvenient in the mine, and an instrument called a hygrometer is used.<sup>1</sup>

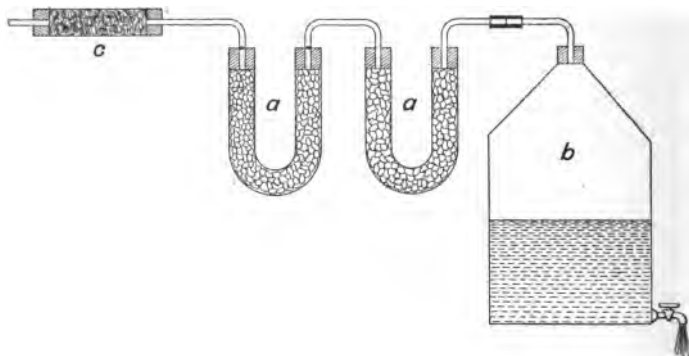
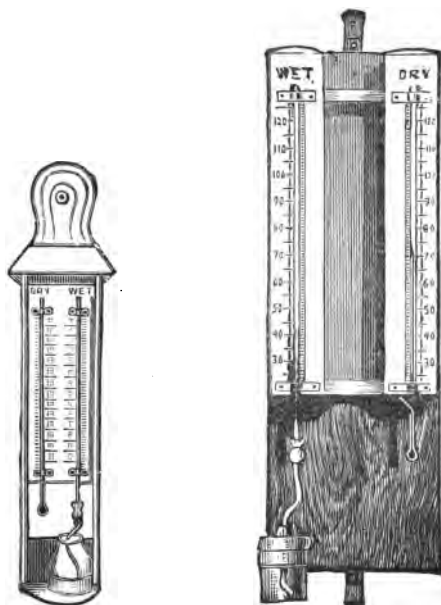


FIG. 20.—Apparatus for moisture estimation.

The apparatus illustrated in Figs. 21, 22 is known as Mason's wet and dry bulb hygrometer and consists of two thermometers placed side by side on a stand, so that the air can circulate freely about them: round the bulb of one of them is fastened a piece of absorbent cotton wick, the ends of which dip in a small vessel of water. The water is drawn up the threads by capillary action, and spread over the surface of the thermometer bulb by means of the wick. Evaporation takes place from the surface, more or less rapidly according to the amount of watery vapour in the atmosphere of the

<sup>1</sup> This should not be confused with the *hydrometer*, an instrument used for measuring the density of liquids.

mine. The more vapour present, the more slowly the water evaporates from the thermometer bulb, and consequently the *less* the cooling effect upon it; the less vapour present in the air, the more rapid the evaporation from the surface of the wick on the thermometer bulb, and consequently the *greater* the cooling effect on it.



FIGS. 21 and 22.—Mason's hygrometer.

Evaporation is always accompanied by loss of heat. It is evident, then, that the *greater* the difference between the readings of the two thermometers, the *less* aqueous or watery vapour present in the air of the mine. Tables have been constructed showing the dampness of the air from its temperature, and from the difference in the two thermometer indications.

Another form of hygrometer is that known as Daniell's (Fig. 23). It consists of a tube bent twice at right angles,

and having a bulb at each end. One bulb is of black glass, and in it is a thermometer and a little ether (a liquid which very readily evaporates). The other bulb is of clear glass, and has a piece of muslin wrapped round it. When ether is poured on this muslin, rapid evaporation takes place, producing cold, and thus causes the ether in the black bulb to evaporate also, and consequently its temperature is lowered.

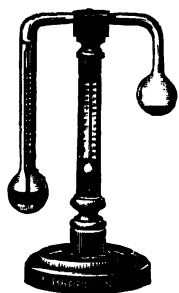


FIG. 23.—Daniell's hygrometer.

Moisture or dew then begins to deposit on the outer surface of the black bulb. When this is noticed the thermometer is read, and the instrument left to itself until the dew disappears again; the thermometer is again read, and the average of the two results taken as the temperature of the dew-point. This is an indication of the amount of moisture in the air; the lower the temperature necessary to

produce dew on the bulb, the less the amount of water-vapour in the air.

The amount of vapour which will saturate a cubic foot of air under standard pressure of the barometer (760 millimetres, practically 30 inches) is as follows:—

At 32° F.	.	.	.	about 2 grains
50° F.	.	.	.	„ 4 „
60° F.	.	.	.	„ 6 „
70° F.	.	.	.	„ 8 „

and a difference of about 5° between the readings of the wet and dry bulb thermometers is equal to about 75 per cent. of saturation, which is held to be a healthy amount.

## CHAPTER IV.

### COMPOSITION OF THE ATMOSPHERE.

THE ancients observed that when certain metals were exposed to the air they lost their lustre and brilliancy, and became coated with a film of something different from the metal altogether; and further, they proved that this alteration was accompanied by an increase in weight.

**Oxygen and Nitrogen.**—Priestley and Lavoisier investigated the matter still more fully, and proved by using mercury, that the film at first produced on the metal when heated in air, could be again decomposed, on further heating, into the metal mercury and a gaseous substance which had evidently been derived from the atmosphere, and is now known as oxygen. It was also proved that a certain amount of gas remained unaltered by the mercury, and incapable of forming any further film of oxide, and thus the air was shown to consist of two gases which came to be called oxygen and nitrogen. Modern investigations have shown that there are other constituents also, in small quantities, to which the names argon and helium have been given, but they are of but passing interest in these notes.

An interesting experiment to show the composition of air is made as follows:—Take a small piece of candle and fasten it on a cork, or in a small porcelain basin, and float it in a shallow tray of water. Now light the candle and place a bell-jar over it; the candle burns with a gradually decreasing flame and finally goes out, and the water will be found to have risen in the jar. The oxygen of the

air has been used up by the burning candle until there was none left to support the combustion, when the candle flame was necessarily extinguished.

A more accurate way of showing the composition of the atmosphere, is to take a measured volume in a glass tube known as a eudiometer, which is sealed at one end and has divisions marked and numbered on it (see Fig. 24). Fasten a piece of phosphorus (which should be handled with great care *under water* because of its inflammable nature) on the end of a piece of wire bent as shown, and introduce it into the air space, the volume of which was previously noted, and leave it thus for a few hours, having a few inches depth of water in the outer vessel. The phosphorus takes up the oxygen from the air, forming a substance (oxide of phosphorus) which readily dissolves in the water. Finally, nothing remains but nitrogen, the volume of which can be measured.

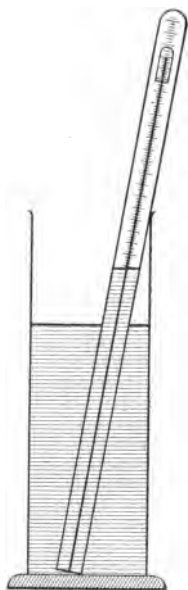


FIG. 24.—Apparatus for testing the composition of the air.

In an actual experiment, the air measured 250 volumes before putting in the phosphorus, and, after the action was finished, the residue measured 198 volumes, which calculates out to 20·8 per cent. of oxygen.

The average composition of *pure dry* air is practically 21 volumes of oxygen gas mixed with 79 volumes of nitrogen gas, and it has been found to vary only within narrow limits in all parts of the globe, and at various altitudes. The air of crowded rooms, theatres, etc. has naturally less oxygen, and the same remark applies to ill-ventilated mines, the oxygen in which may fall almost one per cent below the normal.

If the gases oxygen and nitrogen, separately prepared, be mixed in the above proportions, the mixture possesses

the same properties as ordinary air. When chemical action takes place between the nitrogen and the oxygen of the air, as caused by an electric discharge through them, a reddish-brown gas is produced in which no life could exist. The properties of these two gases are now to be considered.

**Preparation and Properties of Oxygen Gas.**—Oxygen may be made by the method used by Lavoisier and Priestley—namely, by heating mercuric oxide—but the method is not very practicable, and is costly compared with others that are available. The most ready method is as follows (see Fig. 25):—Procure a glass test-

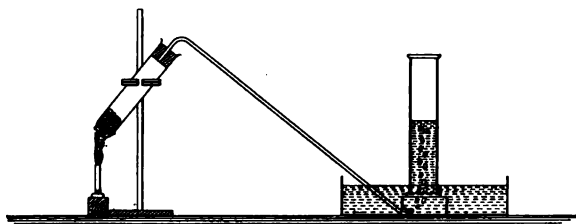


FIG. 25.—Diagram illustrating the preparation of oxygen gas.

tube, about  $6'' \times \frac{3}{4}''$ , and fit it with a good sound cork, through the centre of which a hole has been bored of suitable size to take a piece of bent glass delivery tube. Fill the test-tube about  $\frac{1}{3}$  full of a mixture of potassium chlorate and manganese-di-oxide (4 parts chlorate to 1 part oxide). Now attach to the glass tube fitted through the cork a suitable length of india-rubber tubing, so that the end of the latter may be brought under the shelf of a pneumatic trough, in which an inverted jar is placed filled with water. On applying heat (by means of a Bunsen burner or a spirit lamp) to the contents of the tube, oxygen gas will soon be given off, and will bubble up through the water in the inverted glass jar or cylinder until all the water is displaced, and the jar full of gas. By placing a flat glass plate under the mouth

of the jar, the latter may be removed from the trough, and another one, filled with water, substituted. In this way collect two or three jars full.

By means of a piece of stout wire, bent into a cup shape at one end (called a deflagrating spoon), introduce a piece of burning sulphur into one of the jars. Notice the increased brilliance with which the sulphur burns. Similarly, on placing a piece of glowing charcoal into the gas, it burns with very much increased brightness. A piece of wood, first lit and then partly extinguished, leaving a red-hot spark on the end, on being introduced into the gas at once bursts into flame; and a piece of fine iron wire, tipped with sulphur to start the action, will burn brilliantly in the gas, throwing out sparks of oxide of iron at a white heat.

Several processes have been proposed to obtain pure oxygen from the atmosphere, and one of the most successful of these depends upon the use of a substance—barium oxide—which, when heated in air to a certain degree, takes up oxygen. On increasing the temperature, this combined oxygen is given up again, leaving the original substance to be used over and over again. The gas is largely used in a compressed form, in strong steel cylinders made to withstand a pressure of many atmospheres. One of the most important purposes for which it is used in connection with mining is for restoring a person who has been overcome by noxious gases. In a compressed form it is one of the essential parts of a pneumatophore, an appliance by means of which a person may enter and exist in a noxious atmosphere.

Oxygen gas has very powerful chemical properties, and (so far as we know at present) the chief purpose of the nitrogen gas in the air is to dilute or weaken the action of the oxygen. If breathed alone for any appreciable time it produces feverish symptoms.

In every 18 lbs. of water there are 16 lbs. of oxygen, and almost one-half of the weight of the rocks composing

the earth's crust consists of oxygen, combined with silica, alumina, iron, lime, magnesia, potash and soda, in varying amounts. Most of the chemical actions brought about by the atmosphere are due to its oxygen. Instances are found in the rusting of metals, the breathing of men and animals, the decay and decomposition of animal and vegetable matter—all being processes in which oxygen plays a most important part.

Oxygen is the only gas which will support respiration, and unless it be present *in the free state*, respiration and life are not possible. Oxygen gas is soluble in water, a matter of vital importance to fish life, it being as necessary to their existence as to ours. Hence the need of aerating the water in aquaria. Fish cannot live in water from which the air has been expelled by boiling, or in which the oxygen has been used up in any way.

**Nitrogen and its Properties.**—The other constituent of the atmosphere, nitrogen, is the exact opposite, in chemical properties, to its companion oxygen. Though entering into the composition of certain very active chemical reagents (*e.g.* nitric acid, ammonia, etc.), it is itself very inert, or inactive, chemically, and most of its properties are negative. We have seen that it is widely distributed in Nature, forming (as it does) four-fifths of the atmosphere by volume. It enters largely into the composition of plants and animals, some plants having the power of absorbing this atmospheric nitrogen by means of bacteria on their roots.

Nitrogen will neither burn itself, nor allow anything to burn in it, and it will not support life. This latter is an important fact to remember, because, as will be shown later, *nitrogen is a large constituent of after-damp*. Nitrogen is only very slightly soluble in water, and combines directly with very few substances. The gas may be made by depriving air of its oxygen, as already shown; also by several other processes, which are too purely chemical to describe here.

## CHAPTER V.

### SUNDRY CONSTITUENTS OF THE ATMOSPHERE.

THE composition of *pure dry air* has been previously given, but air in this condition is a rarity, and under ordinary circumstances it always contains, besides oxygen and nitrogen, more or less of the following compounds:—

Aqueous (watery) vapour.

Carbon-di-oxide (commonly called carbonic acid).

Ozone.

Nitric acid.

Ammonia.

Sulphur compounds.

It is well to consider why and how these substances are always present in variable amounts.

**Aqueous Vapour.**—The amount of aqueous vapour present in the air depends upon climatic conditions, and upon the temperature. When the air is holding as much vapour as it possibly can, it is said to be *saturated*, and if when in this condition the air be cooled, some of the vapour will be deposited in the form of dew or rain. There is always more vapour present in the atmosphere in summer than in winter, because the warmer air can hold more in suspension.

The great source of water vapour in the air is from the evaporation from large surfaces of water which is constantly going on. It is a matter of common observation that after a long period of dry weather the level of large sheets of water is perceptibly lower, due to evaporation.

The liquid has gone into the atmosphere, in the form of aqueous vapour. Another source is in the breathing of men and animals, and another in the burning of any substance containing hydrogen (such as the various oils, coal gas, etc.). This latter process is illustrated in Fig. 26, where at B water collects, which is produced from the hydrogen of the coal gas and the oxygen of the air. It is a common experience that if one breathes on a mirror the moisture is condensed on its cool surface; or if a cold surface, say of metal, be held over any gas flame, it becomes wet, because the moisture, produced by the hydrogen of the burning gas, is condensed; and in a crowded room one sees water trickling down the windows and walls especially if the temperature outside be low.

It will be seen from this how useless and insanitary it is to attempt to dry, or *air*, an apartment by lighting gas in the room. The air becomes charged with moisture and also with the impure products of combustion (chiefly carbon-di-oxide, moisture, and sulphur compounds).

**Carbon-di-oxide (Carbonic acid gas).**—This gas is usually present in air to the extent of about .04 per cent. It is being continually produced in all processes of burning carbonaceous matter (*e.g.* coal, wood, oils, etc.); in the putrefaction and decay of organic matter; also by the breathing of men and animals.

By drawing the products of combustion of a gas flame through lime water, as in A, Fig. 26, a milkiess is produced in the latter, proving the presence of carbon-di-oxide. Also by breathing through lime water the same effect is produced, due to the same cause (see also p. 57).

Now since all these sources of carbon-di-oxide are at work, continually discharging the gas into the air, the question naturally arises, What becomes of it all? The answer is that plant-life in general has the remarkable property (in presence of sunlight) of decomposing or splitting up carbon-di-oxide, using the carbon for its

structures, and giving us oxygen back again, for purposes of respiration. This may be readily proved by the following experiment.

Fill a tall narrow bell-jar, having a stopper ground into it, with water, then immerse a good sized bunch of fresh mint, or fresh green leaves of any kind, in the water. Invert the whole in a dish of water and place in the sunlight; bubbles of gas will appear on the leaves, and then detach themselves, collecting at the top of the bell-jar. On applying a glowing chip to this gas, it will be found to be oxygen.

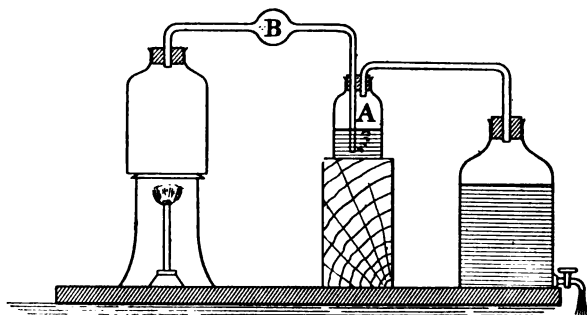


FIG. 26.—Drawing the products of a gas flame through lime-water.

It has been noticed that at high altitudes the amount of carbon-di-oxide increases, no doubt because there is very little plant-life in those regions.

The other properties of carbon-di-oxide are considered later on (p. 56).

**Ammonia** is a compound of nitrogen gas and hydrogen gas, and is produced naturally by the decomposition of bodies containing nitrogen. Its pungent smell can be readily recognised on entering a stable, the door of which has been closed for some little time. It is more likely to be found in air near towns than in fresh country air. The

small amount present in air contributes to plant-life, on account of its nitrogen.

**Nitric Acid.**—This is formed in the air whenever a thunderstorm is in progress, and rain water examined after a storm of that kind is always richer in nitrates than ordinary rain water. These nitrates are a valuable plant food. In some parts of the world, notably in Chili (South America) large deposits of nitrates occur which have been formed by bacterial action (or nitrification).

**Ozone** is never found in the neighbourhood of towns, but only in the freshest country air, and in sea air. It is easily destroyed by decaying or putrefying organic substances of all kinds. The amount varies in country air, but is fairly constant in sea air. It is produced by electrical discharges in the atmosphere, and also (it is supposed) by the evaporation of sea water. In many of its properties it resembles hydrogen peroxide, and no doubt it is often mistaken for this substance.

**Sulphur compounds** are chiefly found in the neighbourhood of towns or works, and are due to the burning of fuel or other substances containing sulphur. Certain sulphur compounds are also due to the decay of animal and vegetable matter containing sulphur (see sulphuretted hydrogen or stink damp).

Sulphur or brimstone itself is found in the free state or *native* condition, naturally, and has therefore been known from the earliest times. It is generally found in volcanic districts, and there are mines of sulphur near Sicily. Many of the ores of metals are compounds of sulphur, as for instance, iron pyrites (bisulphide of iron); galena (sulphide of lead); copper pyrites (sulphide of copper and iron).

When sulphur is heated in presence of air, it burns and forms sulphur-di-oxide, a very pungent smelling gas; but

if heated out of contact with air or in closed vessels it goes off in the form of vapour, and then recondenses on the first cold surface it comes in contact with. This action is called *sublimation*, and is a similar process to the *distillation* of a liquid.

Sulphur is like carbon in the respect of having *allotropic* modifications (see under "Carbon"). That is to say, it may exist in several different forms, thus—rhombic and prismatic sulphur (so-called because of the crystalline shape), plastic sulphur (so-called because of its being able to be moulded like clay in the fingers), and amorphous (without shape) sulphur, are all varieties of one and the same element. Sulphur, like carbon, forms two compounds with oxygen—namely, sulphur-di-oxide and sulphur-tri-oxide, the former already mentioned, the latter forming when combined with water the well-known oil of vitriol or sulphuric acid.

## CHAPTER VI.

### WATER.

**Water**<sup>1</sup> was formerly classed among the elements, but is now readily proved to be a compound of hydrogen and oxygen. By passing an electric current through water,

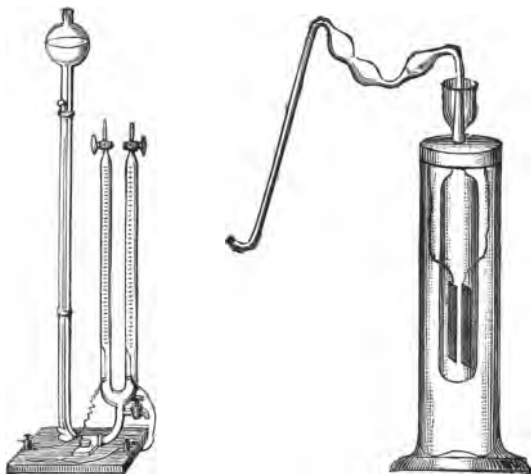


FIG. 27.—Apparatus for decomposing water.

the latter is decomposed, and the two gases which compose it are collected in the two tubes of the voltmeter, as the instrument is called. It is found that the volumes of the gases collected are in the proportion of

<sup>1</sup> Chemical formula  $\text{H}_2\text{O}$ . Density (in the form of steam), 9. Composed of 2 parts by weight of Hydrogen combined with 16 parts by weight of Oxygen. By volume, as water vapour or steam, it contains 2 volumes of Hydrogen combined with 1 volume of Oxygen.

two to one, and that the gas with the larger volume is hydrogen. By taking the gases thus produced into one vessel, and passing an electric spark through them, they re-combine, with an explosion, forming water again.

Another method of proof is to use some substance having a greater affinity or attraction for oxygen than hydrogen has, and which is therefore able to decompose water. Such a substance is found in the metals potassium and sodium, which on being thrown on to water at once decompose it, causing hydrogen gas to be liberated. By enclosing the metal in a kind of gauze spoon, and holding a cylinder filled with water over the bubbles of gas, they may be collected and proved to be hydrogen.

Further, if steam be passed through a red-hot iron

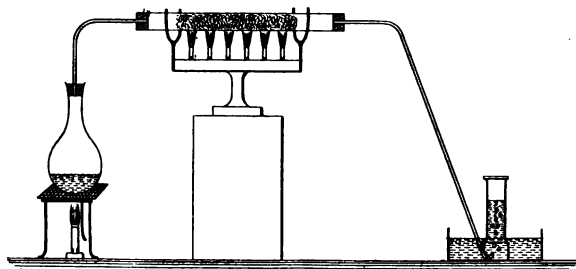


FIG. 28.—To illustrate decomposition of steam.

pipe, it (the steam) is decomposed into its elements, the oxygen combining with the iron and forming iron oxide (magnetic), and the hydrogen gas passing on in the free condition. Thus  $3 \text{ Fe} + 4 \text{ H}_2\text{O} = \text{Fe}_3\text{O}_4 + 4 \text{ H}_2$ . It may be readily collected over water, and proved to answer to all the tests for hydrogen gas (see Fig. 28).

When water is cooled down it obeys the usual law of contraction *up to a certain point*,  $4^\circ \text{ C.}$ , but at that temperature it begins to expand until the freezing point  $0^\circ \text{ C.}$  is reached. (Hence at  $4^\circ \text{ C.}$  water has its maximum density.) This expansion is responsible for the bursting of water pipes,

and for the breaking down of masses of rock and the crumbling of soils. The water, after filling the crevices of the rocks, expands when frozen with an irresistible force.

Bomb-shells have been burst by filling them with water and then exposing them to the temperature of freezing point, or below it.

**Pure Water, and Natural Waters.**—Pure water is a product of the laboratory, for water is never found absolutely pure in Nature. The nearest approach to it is *rain water*, which is really produced by a process of natural *distillation*. The heat of the sun evaporates water from the sea and other sources; it is carried into the atmosphere as vapour or moisture, and is condensed by cool winds, falling again as rain. But the rain water becomes contaminated by impurities in the air through which it has passed, especially in the neighbourhood of towns.

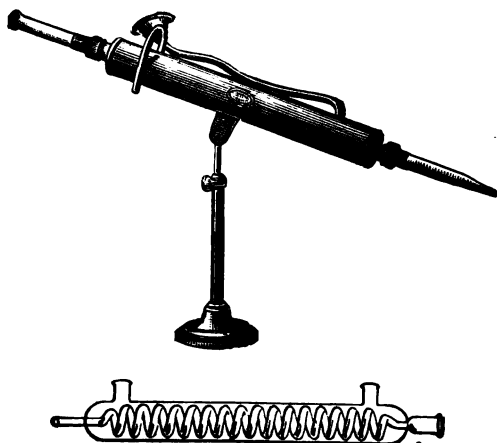
Water is generally purified by distillation, which consists in converting the liquid into steam and condensing the latter by suitable means, generally by passing through a tube or pipe, round the outside of which cold water circulates (see Figs. 29, 30). By these means the solid impurities are left behind in the boiler or retort.

Filtration, which is resorted to as a purifier of water, does *not* remove matter which is actually *in solution*, or dissolved in the water, but only that which is *in suspension* in the liquid. For the supply of towns, water is filtered through beds of sand and gravel. Since water is the most universal solvent we know of, it follows that all water that has been in contact with the earth or other solid matter is liable to be contaminated by various substances. Most river waters are turbid or muddy, owing to the presence of finely-divided solid matter *in suspension*. This would settle out if allowed to do so, but the clear water would still contain more or less solid matter *dissolved* in it. The dissolved substances have been derived from the land, the rocks and strata, over and through which the water

has passed, and they vary in nature and in amount with the kind of rocks traversed. Water passing over magnesian limestone contains, roughly, ten times as much solid matter in solution as that which passes over a granite formation.

Not only is there the solvent action of the water itself, but the latter in passing through the air takes up carbonic acid, which enables the water to dissolve carbonates from any deposits containing them; and thus in many waters are found lime and magnesia carbonates and sulphates, to which substances the *hardness* of water is due.

#### Temporary and Permanent Hardness of Water.—



FIGS. 29, 30.—Condensers.

The hardness caused by lime carbonate and magnesia carbonate can be removed by boiling, or by the addition of milk of lime, the reason being that the carbonic acid, by whose agency they are held in solution, is either driven off (by boiling) or neutralised (by milk of lime). Such hardness is called *temporary*.

The sulphates of lime and magnesia do not owe their presence to the action of carbonic acid, but simply to the solvent action of the water itself. Consequently they are not precipitated on boiling, and are said to cause *permanent* hardness.

Water may be sometimes so highly charged with carbonate of lime that objects placed in it become, in time, coated with a layer of it, causing *petrification*. In most cases the shells of shell-fish are built up of this substance, which has been assimilated from the water by their organisms. Again, water charged with carbonate of lime and magnesia, in dripping from the roof of a cavern, deposits lime by degrees, forming a growth from the roof known as a *stalactite*, and a corresponding growth on the floor where the water drops called a *stalagmite*; eventually the two meet, and form a sort of pillar of lime and magnesia carbonates.

Certain springs—for instance, those at Epsom, in Surrey—are charged with magnesium sulphate (hence the name Epsom salts); others (such as the chalybeate springs found in various parts of the country) contain salts of iron, etc.; and the writer had a particular case under examination some little time ago, where very decided amounts of manganese (*not* magnesia) were found in solution in a natural water, probably as bicarbonate.<sup>1</sup>

**Water for Boiler Purposes.**—The importance of these facts to persons identified with mining lies in the circumstance that all waters, which have to be dealt with in mines, are contaminated, more or less, by salts of different kinds and in varying quantities; and upon the nature and amount of these solids it depends whether a given water can be used for boiler purposes. This is a matter which can only be thoroughly decided by analytical examination. Not only should the total solids be known, but also the nature of those solids, and the hardness of

<sup>1</sup> See *Chemical News*, October 1901.

the water. Further, it is necessary to know whether the water is acid, or not.

A water containing, say, 60 or 70 grains per gallon of sulphate of soda, or carbonate of soda, would not be nearly so objectionable as one containing 8 or 10 grains per gallon of carbonate or sulphate of lime or magnesia. Sulphate of lime is a most objectionable constituent, as it forms such a hard, compact scale on the boiler plates.

The following are analyses of boiler scales, which have been formed in boilers fed with waters from various districts:—

	No. 1. p. cent.	No. 2. p. cent.	No. 3. p. cent.	No. 4. p. cent.	No. 5. p. cent.	No. 6. p. cent.
Silica (sandy matter) . . . }	1·50	2·50	5·20	5·60	9·10	5·0
Iron oxide and alumina (clayey matter) . . . }	·50	·50	4·8	1·60	6·50	2·0
Lime carbonate . . .	—	traces	70·6	82·1	40·0	73·2
Lime sulphate . . .	—	82·5	·87	trace	11·6	traces
Magnesia carbonate . . . }	98·0	14·5	18·1	10·4	29·2	20·0

Waters containing magnesium chloride are to be avoided, as at high temperatures hydrochloric acid is liable to be produced, which, of course, would corrode the boiler plates.

A *thin* scale in a boiler is not undesirable, as it tends to neutralise, or destroy, any acids that may be formed. The water from a peaty moorland is often found to be acid, and many pit waters are very decidedly so with sulphuric acid, due to the weathering of *coal-brasses*, or iron pyrites, a compound of iron and sulphur. Ochrey waters are often acid.

When the use of waters containing lime and magnesium salts is unavoidable, methods are often made use of for the removal of these objectionable scale-forming substances. These methods are many and varied. The principle underlying some is to precipitate

or throw down the lime and magnesia salts, by means of suitable chemicals, generally in the form of carbonate of soda, caustic soda, phosphate of soda, milk of lime, used singly, or mixed, to suit the particular needs of the water to be treated. The precipitated matter is allowed to settle in tanks, and the clear liquid is siphoned or pumped off; or the sediment is separated by passing through a series of partitions in which are baffle plates and filtering medium, such as wood pulp, shavings, or coarse matting.

Most of these processes, as a rule, give good results, provided that they have proper attention, and that the supply of chemicals be properly regulated according to circumstances. After a wet period most water supplies require *less* chemicals per 1000 gallons than they do in a dry period, as in the latter case the solids in the water are getting concentrated. After treatment, the water is said to be *softened*. It is, as a rule, preferable to treat the water *before* going into the boiler, when this is practicable.

Besides the chemical reagents above mentioned, other substances are used, such as tannate of soda, argol (a crude tartrate of potash), petroleum, and tannery refuse. Zinc plates are, in some cases, suspended in the boiler, the galvanic action set up between the zinc and the iron being said to have a beneficial effect in scale prevention. Oyster shells and even potatoes are used in many cases. Greases are, as a rule, to be avoided, as they are liable to give rise to acid substances at the temperature which exists in the steam boiler.

The pitting or corrosion of boilers may be due to the presence of acid (sulphuric acid as before mentioned), also of hydrochloric acid from magnesium chloride. It is also thought possible that *local action* may cause it in some cases—that is, a slight difference in chemical composition between two adjacent plates, giving rise to galvanic action; or a difference brought about in punching

or shearing the plates, whereby the material near the punched hole is altered somewhat, physically, from the rest of the plate, again giving rise to galvanic action.

The *external corrosion* of boilers is sometimes a cause of weakening. This is, of course, more likely to occur when the boiler is out of use. When in use, the sulphurous acid gas from the pyrites of the coal is an active agent of corrosion on the parts of the boiler farthest removed from the fire, upon which dampness or moisture may condense at some time. Soot and dirt on these parts encourage the retention of water, and therefore increase the tendency to corrosion, since they help to retain the acid fumes. Hence the boiler should be inspected regularly, and kept as clean and dry outside as possible.

**Hydrogen.**—It is important that this gas, which with oxygen forms water, should have here a description of its properties and methods of formation.

A curious and interesting sight is to see the liberation of hydrogen from water by means of the metals potassium and sodium. The water appears to take fire owing to the flame produced due to the ignition of the liberated hydrogen.

The most general method of making hydrogen is by decomposing an acid, such as hydrochloric or sulphuric acid, by means of the metals zinc or iron. Fig. 31 shows the apparatus used for this purpose. The zinc is placed in the flask, fitted, as shown, with a safety funnel, down which the diluted acid is poured. The gas begins to come off immediately, but time must be allowed for the air in the flask to be expelled before collecting. It is collected over water in a glass cylinder placed on a beehive shelf, as shown.

Hydrogen is the lightest substance known, being about  $14\frac{1}{2}$  times lighter than air. On this account it is used for filling navigable balloons, etc., though in some cases coal-gas is used for cheapness' sake. It may be collected in a

vessel held upside down, and poured *upwards* from one vessel to another. It is supposed to be the vapour of a very volatile metal to which the name hydrogenium has been given. Certain metals, for instance platinum, palladium, and even iron, absorb the gas or *occlude* it, and many useful applications are based upon this property with regard to the first two metals above named.

Those interesting objects known as *meteorites*, which fall from time to time on our sphere, often contain

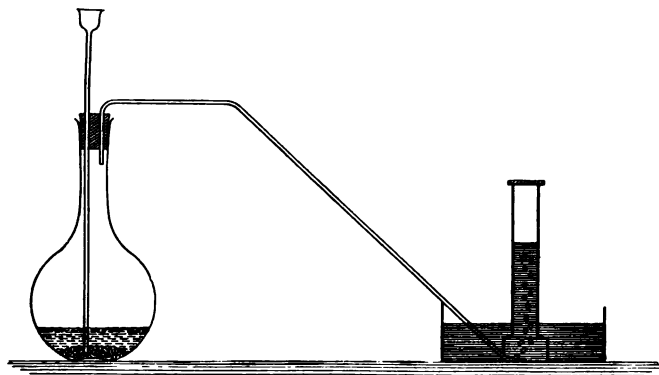


FIG. 31.—Apparatus for the preparation of hydrogen gas.

*occluded* hydrogen gas, and we know beyond doubt (by means of spectroscopic observations) that the atmosphere surrounding the sun (more correctly called the *photosphere*) is one huge volume of hydrogen gas in an incandescent state.

Hydrogen is necessary to the constitution of all true acids, though many substances are commonly called acid which do not contain hydrogen (notably carbon-di-oxide).

One-fourth of the weight of marsh gas or fire-damp is due to hydrogen. It is very inflammable, and forms a highly explosive mixture with air or oxygen. When it burns, water is produced. The gas has no poisonous

or shearing the plates, whereby the material near the punched hole is altered somewhat, physically, from the rest of the plate, again giving rise to galvanic action.

The *external corrosion* of boilers is sometimes a cause of weakening. This is, of course, more likely to occur when the boiler is out of use. When in use, the sulphurous acid gas from the pyrites of the coal is an active agent of corrosion on the parts of the boiler farthest removed from the fire, upon which dampness or moisture may condense at some time. Soot and dirt on these parts encourage the retention of water, and therefore increase the tendency to corrosion, since they help to retain the acid fumes. Hence the boiler should be inspected regularly, and kept as clean and dry outside as possible.

**Hydrogen.**—It is important that this gas, which with oxygen forms water, should have here a description of its properties and methods of formation.

A curious and interesting sight is to see the liberation of hydrogen from water by means of the metals potassium and sodium. The water appears to take fire owing to the flame produced due to the ignition of the liberated hydrogen.

The most general method of making hydrogen is by decomposing an acid, such as hydrochloric or sulphuric acid, by means of the metals zinc or iron. Fig. 31 shows the apparatus used for this purpose. The zinc is in the flask, fitted, as shown, with a safety funnel, into which the diluted acid is poured. The gas comes off immediately, but time must be allowed for the air in the flask to be expelled before the gas is collected over water in a glass cylinder inverted on a shelf, as shown.

Hydrogen is the lightest substance known, being  $14\frac{1}{2}$  times lighter than air. On this account it is used for filling navigable balloons, etc., and for cheapness of gas is used for cheapness.

## WATER

vessel held upside down, and poured into another. It is supposed to be a very volatile metal to which the most attention has been given. Certain metals, such as radium, and even iron, absorb the gas in many useful applications. In fact, with regard to the first two metals, they

Those interesting objects which fall from time to time from the sky

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is harmless.  
contains  $\frac{3}{4}$  of its  
oxide contains  
le),  $\frac{3}{11}$  of its  
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carbon is a large  
of his flesh and

and calcium, in  
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carbon on steel is very  
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it is peculiar in that it

exists as carbon *alone* in very different forms (spoken of as *allotropic* modifications). For instance, the diamond, graphite (from which pencils are made), and charcoal, can be easily proved to be *different forms* of the *same element*, carbon. When burnt in oxygen gas they each form the same compound, carbon-di-oxide. A weighed quantity of diamond produces the same quantity of gas as the same weight of pure charcoal does. Other common forms of carbon are lampblack, gas-carbon, animal charcoal, wood-charcoal, etc.

**Lampblack** is made by burning substances rich in carbon, such as oils, tar, tallow, etc., with as little air supply as possible; and the smoke from them, which is chiefly carbon in a very fine state of division, is condensed in some suitable way in cool chambers, from which it is collected afterwards.

**Gas-Carbon**, as its name implies, is a product of the gasworks, and has been slowly deposited on the sides of the retorts in which the coal is distilled for gas-making purposes. It is a very pure form of carbon, for a commercial product, and contains little ash. It is used for making carbon rods for electric lighting and for electric battery purposes, also for recarburising, or adding carbon to molten steel.

**Coke** is a much more impure form of carbon, and contains all the mineral matter which was originally present in the coal from which it was made. It contains, on an average, about 88 to 90 per cent. carbon (see under "Fuels").

**Animal Charcoal, or Bone Black**, is made by heating bones in vessels from which air is excluded. It contains very little carbon. About 90 per cent. of it is phosphate of lime, but owing to its peculiar condition it forms a sort of carbon sponge. It is used for clarifying certain liquids,

*e.g.* in sugar refining, and for various other purposes of decolourising.

**Ordinary Charcoal** is made by heating wood in closed vessels, no air being allowed to enter. It bears the same relationship to wood that coke does to coal, and retains all the mineral matter originally present in the wood, this forming the ash. Charcoal has the remarkable property of absorbing gases in its pores, and on this account is able to prevent certain foul gases becoming injurious. A layer of charcoal spread over decaying or decomposing matter of any sort will at once arrest the smell.

Charcoal is added to soils, in some cases in order to keep them sweet and pure, and trays containing this substance are often placed in the ventilators over sewer pipes to absorb ill-smelling vapours. The reason why charcoal and porous cinders will float on water is because of the air contained in their pores. If this be removed under an air pump they will sink.

The compounds of carbon and oxygen next claim attention. There are two, namely, carbon-mon-oxide, and carbon-di-oxide.

**Carbon-mon-Oxide, Carbonic Oxide, or White-Damp.**—Chemical formula CO. Density 14. Specific gravity .967. It is composed of 12 parts by weight of carbon combined with 16 parts by weight of oxygen; by volume it contains one volume of carbon vapour combined with one volume of oxygen. This is the most treacherous gas with which the miner has to deal. There are no ready means of detecting its presence. It has no colour, no taste, no distinct smell; yet the presence of so small a quantity as 1 *per cent.*, if breathed for any length of time, is sufficient to cause death. It cannot be too forcibly realised that this gas is as much a *poison* as prussic acid and arsenic, and does not kill by *suffocation*, or because of the deprivation of oxygen, as some gases do.

The following effects are produced when carbonic oxide is breathed :—

Percentage of gas.	Effect.
·05	After half an hour or more, slight giddiness on exertion.
·10	After half an hour or more, inability to walk.
·20	After half an hour or more, loss of consciousness, and perhaps final death.
1·00	After a few minutes, loss of consciousness, and final death.

Oxygen gas is the best antidote to administer to a person who has been “gassed” with carbonic oxide.

In the pure state the gas will burn, forming the other oxide ( $\text{CO}_2$ ), and in certain proportions it forms an explosive mixture with air. The most explosive mixture is two volumes of CO with five volumes of air. This gas can be often seen burning with a blue flame on the top of a caked coal fire, and is always produced under similar conditions to these—namely, when carbon in any form burns with a limited or deficient supply of air. This may be experimentally proved by passing the di-oxide through a red-hot tube packed with charcoal (using the same apparatus as shown in Fig. 26), packing the tube with charcoal instead of iron; carbon-mon-oxide issues from the other end of the tube. In the fire-grate the gas  $\text{CO}_2$  is formed at first, but on passing up through the hot mass of coal it takes up more carbon, and becomes carbonic oxide (CO)—thus  $\text{CO}_2 + \text{C} = 2\text{CO}$ .

Dense black smoke is, as a rule, accompanied by carbonic oxide, and this means a *loss of fuel* in two ways—namely, the carbon, forming the black particles to which the smoke owes its colour; and the carbonic oxide, which is *unburnt* gas. Both of these substances, if burnt, would give out their proportion of heat.

**White-Damp** is not found naturally in the air of the

mine, but is formed in the explosion of certain compounds used for breaking down the coal. It is also produced by *gob fires*, in which carbon, in the form of coal, is burning with a limited supply of air, and is present in *after-damp*—that is, the mixture of gases formed by an explosion of *fire-damp* in the mine. In many cases, those who have escaped the direct effects of an explosion have been overcome and poisoned by this gas. The suicide, who seals himself up in a room in which a charcoal stove is burning, owes his death to white-damp. The gas from blast furnaces used in iron smelting contains, roughly, 30 per cent. of white-damp, which was formerly allowed to escape into the atmosphere, but is now conducted under boilers and there burnt, and its heat of combustion made use of.

No very ready test for the presence of carbonic oxide is available, especially for recognising such small amounts as may be fatal to life. If present in any quantity, the cuprous chloride reaction may be used, but this is only the case where the actual amount is to be estimated. The gas is soluble in this reagent (cuprous chloride solution).

The most practicable method of recognising the presence of small (or large) quantities of the gas depends on the action of the latter upon the blood. If a few drops of blood be added to water a pale yellow colour is produced, and if this be shaken up with carbonic oxide gas the colour changes to pink. The effect on blood is also readily seen by means of the spectroscope, an instrument, however, which could scarcely be used under the requisite conditions for testing a noxious atmosphere in the mine.

Dr Haldane suggests the use of a mouse or a bird, by noticing the effects on which creatures the safety (or otherwise) of an atmosphere can be gauged. The effect of carbonic oxide on these animals is much quicker than on a human being, and thus gives timely warning.

**Carbon-di-oxide, Black-Damp, Choke-Damp or Stythe**, commonly known as carbonic acid gas. Chem-

ical formula  $\text{CO}_2$ . Density 22, specific gravity 1.527. It contains 12 parts by weight of carbon combined with 32 parts by weight of oxygen, or one volume carbon vapour combined with two volumes of oxygen.

In the atmosphere of the mine this gas is always present in greater or less quantity, being produced continually by the breathing of men and animals, and also by the burning of the lamps; and it is often given off from the coal itself. It is always found in *after-damp*, having been formed by the explosion of marsh gas or fire-damp, a substance rich in carbon. Carbon-di-oxide is a colourless invisible gas, with a *very* slightly acid taste and smell, but it is rarely in sufficient amount to be recognised by these properties: in fact, such an atmosphere would not support life. It will not burn itself, nor allow anything to burn in it. The sparkling of aerated waters is due to the escape of this gas, which has been forced into the water under pressure. Its weight enables it to be poured, siphoned, or ladled from one vessel to another, as though it were a liquid, since it is  $1\frac{1}{2}$  times as heavy as air.

By balancing a glass vessel at one end of a chemical balance, and then displacing the air in it by carbon-di-oxide gas, the increase in weight is readily seen.

The property that this gas has of extinguishing a light has been made use of in fire-extincteurs, the constituents of which give out carbon-di-oxide. It is also used in a compressed form in strong cylinders for the same purpose, and has been made use of for extinguishing fires in mines. Large tanks are built at the surface, pipes from which lead the heavy gas into the workings. The gas is generated by the action of hydrochloric acid on limestone or chalk.

It must not be supposed that, on account of its weight, the carbonic acid is *always* found on the *floor* of a mine (though it is often found lying in holes or pockets in the ground and in lower level workings). This would be so were it not for *gaseous diffusion*, which has been already explained (Chapter III.).

In some cases, however, such a phenomenon is to be noted, as in wells, and in certain districts where the gas issues in continuous discharge from the earth (from subterranean sources), notably in the Poison Valley of Java, where a kind of lake of the gas exists, fed from the earth. In such cases the safety of the atmosphere for breathing is readily tested by lowering a lighted candle into it. If it continues to burn, then carbon-di-oxide cannot be present in dangerous quantity. It is possible to exist for a while in an atmosphere which will not support the combustion of a candle, but it is extremely dangerous. When the amount of carbon-di-oxide gas reaches 8 to 10 per cent., the atmosphere will not support life long.

In a paper read before the Society of Chemical Industry, in April 1895, by Clowes and Feillman, it was pointed out that there was a marked similarity in the composition of the gases produced by breathing, and by combustion. Thus air from the lungs, after the breath had been held for forty seconds, gave on analysis—

14.9% oxygen; 81.4% nitrogen; 3.7% carbon-di-oxide.

The gases produced by the burning of an alcohol flame contained—

14.9% oxygen; 80.7% nitrogen; 4.3% carbon-di-oxide;  
and from a burning candle—

15.7% oxygen; 81.1% nitrogen; 3.2% carbon-di-oxide.

It was stated, on the authority of Dr Haldane, that atmospheres such as the above did not produce any noticeable inconvenience when breathed. It therefore follows that atmospheres which extinguish any flame ordinarily used *may* yet be perfectly safe to breathe. This does not apply to a hydrogen flame, which is only extinguished in an atmosphere which is quite incapable of supporting life—namely, 5.5 per cent. oxygen and 94.5 per cent. nitrogen.

The following table shows the results from the presence of the stated percentages of carbon-di-oxide :—

Per cent. Carbon-di-oxide Gas.	Effects on Respiration and Combustion.
3·5	Breathing deepens. Lights still burn.
6·0	Marked panting. Lights still burn.
10·0	Severe distress. Lights still burn.
15·0	Partial loss of consciousness. Lights extinguished.
25·0	Death.

Carbon-di-oxide is readily absorbed by a solution of potassium hydrate (or by the solid itself, commonly known as caustic potash). This property is made use of in pneumatophores, the carbon-di-oxide produced by the breathing of the person wearing the apparatus being taken up by the potash. (For description of pneumatophores, see pp. 96-7).

A ready test for the presence of carbon-di-oxide is found in lime-water, which is made by shaking up freshly-burnt lime, or quicklime, with water, and allowing it to settle, using the clear liquid which remains. On breathing through this clear solution for a few seconds, a milkiness is produced in it, expired air containing about 4 per cent. of the gas  $\text{CO}_2$  under ordinary conditions. The milkiness is produced by this carbon-di-oxide combining with the lime in solution in the lime-water, forming carbonate of lime. If the blowing be continued long the milkiness disappears, but is re-formed on boiling, and in this experiment we imitate the process by which carbonate of lime gets into water, causing temporary hardness (see under "Water").

Carbon-di-oxide does *not* exert a *poisonous* action like carbonic oxide does, but deprives one of life by suffocation, on account of the absence of *free* oxygen. Angus Smith devised a rough test for the amount of carbon-di-oxide in an atmosphere thus: He took a bottle of  $10\frac{1}{2}$  ounces capacity, and added  $\frac{1}{2}$  ounce of clear lime-water, shaking up well. If no milkiness be produced, then the sample

of air in the bottle at the time could not have contained more than .06 per cent. of carbon-di-oxide :—

**Preparation of Carbon-di-Oxide or Black-damp.**

—The ordinary method is to decompose limestone with hydrochloric acid, in an apparatus similar to that used in preparation of hydrogen, but almost any carbonate (for instance, washing soda) may be used, and an acid such as vinegar. Many interesting experiments may be made with it, not requiring much apparatus. It may be generated from the above-named materials in a jug or bottle, and *poured* into other vessels as though it were a liquid. In this way the extinction of flame, the effect on lime-water, and its great density, may be readily shown.

No doubt it appears curious, if not impossible, that an invisible gas such as this should contain the black substance, carbon, but this also may be clearly proved by collecting a jar of gas, and plunging a piece of burning magnesium wire into it. The latter metal has a greater affinity or attraction for the oxygen of the gas, than carbon has; consequently, the gas is decomposed, and carbon in the form of small, black, solid particles is liberated, and is readily seen. The white substance (magnesia, or oxide of magnesium) which is produced at the same time may be separated by adding a little dilute acid, then leaving the black particles of carbon more plainly visible still.

## CHAPTER VIII.

### FIRE DAMP.—COMBUSTION.—COAL DUST AND ITS ACTION.

**Methane, Marsh Gas, Light Carburetted Hydrogen, Fire-damp.**—Chemical formula  $\text{CH}_4$ . Density, 8. Specific gravity .55. Composed of 12 parts by weight of carbon combined with 4 parts by weight of hydrogen, or one volume carbon vapour combined with four volumes of hydrogen.

This gas is one of a numerous class of compounds of carbon with hydrogen, called by the general name of *hydrocarbons*. Of all these, fire-damp is the one which contains the largest amount of hydrogen. It is very like hydrogen in its properties, and though eight times denser than hydrogen, yet compared with air it is an extremely light gas. Because of this lightness, it tends to collect in the highest places in the roof of the mine, unless thoroughly swept out and diluted by the ventilating current, assisted by diffusion.

Fire-damp often goes by the name of marsh gas, because it is the gas liberated when the decomposing vegetable matter or deposit at the bottom of ponds or marshes is stirred up. It is evidently the result, in nature, of the decomposition of vegetable organic matter. In this way it has been generated during the formation of the coal measures, countless ages ago, and after being long pent up escapes from the coal and adjoining strata when these are disturbed. Mixing with the air, it gives rise to the extremely explosive mixture, the ignition of which has caused so many disastrous effects in mines throughout the world. The most explosive mixture is about 9.4 per cent., by volume, of fire-damp, mixed with

90·6 per cent. of air. Fire-damp, along with carbon-di-oxide, is often met with, even in sinking through strata, before the coal is reached.

The gas exerts no poisonous action, but will not support life, because of the absence of *free* oxygen. It will burn itself, but will not allow a light to burn in it.

Though pure methane, chemically prepared, has no smell, fire-damp has a slight one, due to the presence of small amounts of other hydrocarbons. When the gas burns it produces carbon-di-oxide and water vapour—the same products, it will be noticed, as in ordinary breathing—and when fire-damp explodes, the products are the same, since, as we shall see, explosion is really a very rapid burning.

#### Gases produced by an Explosion of Fire-damp.—

At this point it will be well to consider what gases are produced by an explosion of fire-damp.

The elements present before the explosion are :—

Carbon	}	constituting fire-damp.
Hydrogen		

Oxygen	}	constituting the air.
Nitrogen		

When the explosion takes place, the carbon of the fire-damp combines with the oxygen of air, producing carbon-di-oxide or black-damp, and possibly some carbonic oxide or white-damp (depending on conditions prevalent at the time of the explosion). The hydrogen of the fire-damp combines with the oxygen of the air to form water, and the *nitrogen gas is left unaltered*.

The latter fact is sometimes lost sight of in considering the gaseous products of an explosion; but, as we have seen, since it cannot support life, it is quite as deadly as black-damp or white-damp. It is also left *in large quantity*, since it forms four-fifths of the volume of the air originally present. So that *after-damp*, as the gaseous

products of an explosion of fire-damp and air are called, consists of a mixture of black-damp, white-damp, nitrogen, and water vapour.

**Laboratory Preparation of Methane.**—Methane may be prepared in the laboratory by heating a mixture of sodium acetate and sodium hydrate, with the addition of some lime to prevent the other materials from melting.

Thus prepared the gas has no distinctive smell, no taste, and burns with a slightly luminous flame (probably due to particles of soda salts suspended in it). Its properties may be well illustrated by means of ordinary coal-gas, which it resembles in composition.

#### **Tests for the Presence of Fire-damp in Mine.**

—The presence of fire-damp is recognised by its peculiar action upon the flame of a safety lamp. Persons having an intimate knowledge of the mine, and of the most likely places for the gas to accumulate, carefully raise a lighted safety lamp into the suspected atmosphere. If the flame suddenly lengthens, it denotes the presence of fire-damp in dangerous quantity; at the same time the flame becomes less luminous and the interior of the lamp cylinder becomes filled with burning gas, giving out a peculiar faint bluish light, known in some districts as the *corpse-light*.

In case the flame does not alter in appearance, there may be still a decided amount of fire-damp present, though perhaps not in dangerous quantity. Under these circumstances a further test is made by means of a very low flame in the lamp, when if fire-damp be present, a sort of cap of pale blue flame surmounts that of the lamp. The quantity of gas that can be recognised in this way may be as low as 2 per cent.

It is pretty generally recognised that the presence of fire-damp is shown on the lower flame of an ordinary Davy lamp when it exceeds 2 per cent., as a pale blue

"cap" or halo, and the flame increases in length until about 6 per cent. of fire-damp is present; from 8 to 15 per cent., the mixture will explode in the lamp; and if the proportion of fire-damp reaches as high as 30 per cent. the lamp flame is extinguished. It has been noted, in considering the properties of hydrogen gas, that it, also, will not support combustion, although burning itself and forming a very explosive mixture with air.

It may be necessary, sometimes, to know the amount of dilution with air required to make a certain volume of fire-damp harmless. The following is an example:—

A mixture of 50,000 cubic feet of air and fire-damp contains 10 per cent. of the latter. How much air must be added to it so as to prevent the "cap" showing in an ordinary safety lamp?

The 50,000 cubic feet of mixed gases must consist of

$$\begin{array}{r} 45,000 \text{ cubic feet air} \\ 5,000 \text{ cubic feet fire-damp} \\ \hline 50,000 \end{array}$$

Assuming that  $2\frac{1}{2}$  per cent. of fire-damp can be recognised by an ordinary safety lamp but that 2 per cent. cannot, it will be necessary to dilute with air so that the fire-damp be present only to this extent (2 per cent.).

The 5000 cubic feet of fire-damp then must be diluted, so as to be  $\frac{2}{5}$  part of the total volume.

$$5000 \times 50 = 250,000$$

But there are already 45,000 cubic feet of air present, therefore the amount to be added is

$$250,000 - 45,000 = 205,000 \text{ cubic feet air.}$$

In the first testing, before any idea of the conditions is known, great care must be taken not to tilt the lamp, but to keep it quite straight (perpendicular): otherwise the lamp flame may heat the gauze of the lamp, possibly to the extent necessary to allow of combustion, and so cause explosion.

**Explanation of the Flame Test for Fire-damp.**—After many careful observations, it has been found that the size of a flame is governed to a great extent by the proportion of oxygen present in the atmosphere in which it burns. If this proportion be *increased*, then the size of the flame is *decreased*: if the proportion of oxygen be *decreased*, then the size of the flame is *increased*. Hence the lengthening of the lamp flame when introduced into an atmosphere containing fire-damp, in which there is less oxygen than in air. It is really as though the combustible gas, whatever it is (oil vapour or hydrocarbons, in the case of a pit lamp), has to search for food in the shape of oxygen, and the less the supply of oxygen the greater need for an *extended surface* of flame.

In a paper read before the British Association in 1892, Dr Clowes described a miner's safety lamp in which an ordinary flame can at will be replaced by a hydrogen flame, and the use of the latter enables the miner to detect readily, and with certainty, the presence of fire-damp when in quantities ranging from .25 to 3.0 per cent., and to measure the amount. The lamp is available both for testing for the presence of gas, and also for illuminating purposes (when the delicate testing for gas is done), by means of an ordinary oil-flame. Hence, the latter alone can be used if necessary for the detection of amounts of fire-damp greater than 3 per cent.

The hydrogen gas is carried in a compressed form in a small strong steel or copper cylinder, and is led into the lamp through a fine metal tube which terminates near the wick (see Fig. 32). Experiments were made to compare the flames of alcohol, benzoline, and hydrogen of equal heights when exposed to an atmosphere containing 1 per cent. of coal gas. In the case of the hydrogen flame the cap was nearly four times as high as that seen over the benzoline flame, and half as high again as that over the alcohol flame. The lamp proved a useful test in atmospheres charged with petroleum or naphtha vapour, which

form an explosive mixture with air.  $\frac{1}{88}$ th of the amount necessary to form an explosive mixture can be detected by the flame.

It was also explained that the amount of inflammable gas in air was measured not only by the size of the flame but also by its appearance, which changes considerably as the percentage of gas rises. Since the presence of coal dust lessens very much the amount of gas necessary for an explosion, a delicate test for the presence of gas is necessary. The presence of a cap over the lamp flame is caused by the increased temperature, enabling a mixture of fire-damp and air to burn (really helping the combustion), although it would not do so at ordinary temperatures. Consequently the *increased* heat, and the *decreased* light of the hydrogen flame, both tend to make the indications of fire-damp more readily seen.

The most recent development of this lamp (Clowes and Redwood's patent) is illustrated in Fig. 32. It is a modified Gray oil lamp. The lamp is charged with a mixture, in equal proportions, of colza and paraffin oils. An upright strip about an inch in breadth of the inside of the lamp glass is smoked dead black by a wax taper, so as to form a background for observing the flame caps. An upright graduated gauge-wire is inserted in the wick holder to read off the heights of the caps.

In testing for gas down to 3 per cent., the sleeve or the pillar of the lamp having been pushed up, the opening thus exposed is closed by the thumb. The lamp, which takes its air supply near the top, is then gradually raised towards the roof, and the behaviour of the flame is noted. If the flame seems likely to be extinguished by gas, the thumb is removed to allow the air-feed to be taken from a lower level, and thus to save the flame. A rough test is thus made in the ordinary way, and serves to detect the larger percentages of gas. If it is wished to measure the percentage of gas present, the oil flame is pulled down by the "pricker" until it just loses its bright tip, and the cap is

looked for against the black smoked glass as a background.

3 per cent. yields a cap of .3 inch.

4   "       "       "       .5   "

5   "       "       "       1.1   "

6 per cent. yields a cap which just reaches the top of the glass. (See Fig. 33.)

In testing for less than 3 per cent., if no distinct cap is seen over the reduced oil-flame, this does not necessarily prove the absence of gas. The hydrogen flame, however, detects and measures gas down to .25 per cent. The cylinder containing the compressed hydrogen serves as a convenient handle to the lamp (see Fig. 33). The hydrogen gas is gradually admitted to the jet in the lamp. As soon as the hydrogen gas becomes ignited the oil flame is extinguished. By turning the valve, the tip of the hydrogen flame is made exactly level with the top of the upright wire, or with the first cross wire if the scale be used. The flame is now of standard height. Careful observation is made to see if any cap is

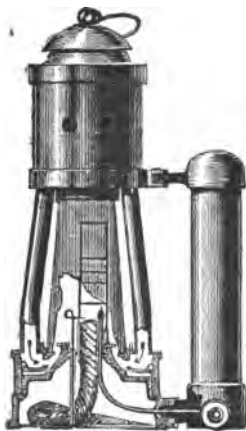


FIG. 32.—Clowes's hydrogen lamp.

visible over the flame. If any cap shows, then the percentage of gas present may be judged by the height and appearance of the "cap."

$\frac{1}{4}$  per cent. of gas gives .7 inch cap (the cap being pale and hazy in outline).

$\frac{1}{2}$    "       "       .7   "       } (more definite in appearance and outline).

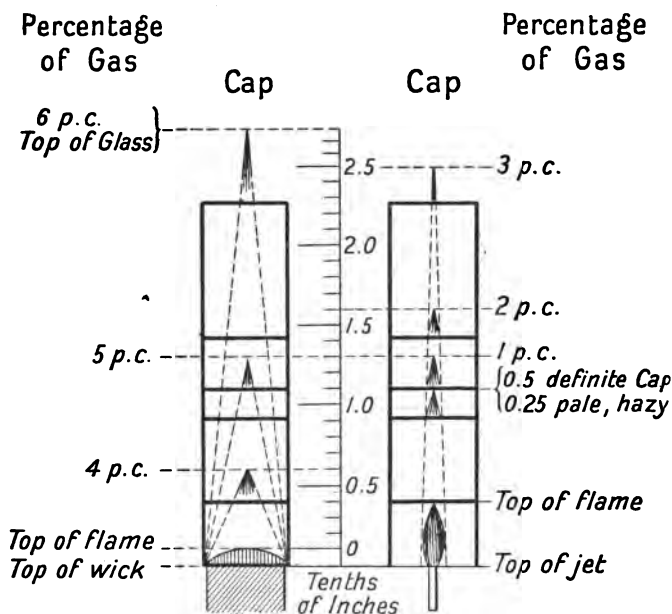
1   "       "       .9   "       } (cap becomes not only larger but also more visible and definite).

2   "       "       1.2   "       }

3   "       "       2.1   "       }

When the test with the hydrogen flame is finished, the wick is pushed up by the pricker and becomes kindled by the hydrogen flame, which is then shut off.

The hydrogen test will probably be applied principally to the "main return," and possibly to "district returns,"



The tip of the cap is 0.2 inch above the horizontal wire except for 6 p.c

FIG. 33.—Illustrating heights of flame "caps."

to ascertain their condition, and to regulate the ventilation current and its distribution accordingly. It will be understood that neither the oil nor hydrogen flame must be extinguished until the other flame is burning satisfactorily in the lamp.

Several lamps of various forms have been devised which will detect as little as  $\frac{1}{4}$  per cent. of fire-damp in an atmo-

sphere. Most of these are only intended to be used *after* an ordinary lamp test has failed to give any indication of fire-damp. The alcohol flame of the Pieler lamp will detect  $\frac{1}{4}$  per cent. of fire-damp readily. The flame gives out no light, and is chiefly used for gas testing. Some lamps use benzoline, and are useful in that they give a fair amount of illumination, as well as testing for presence of gas. As a rule, those in which inflammable liquids are used have a sponge in the oil tank, which soaks up the liquid, any excess of which is drained off.

An ingenious method for indicating fire-damp has also been used by Liveing, based upon the fact that if the metal platinum be warmed, and then placed in an atmosphere of hydrogen, or of coal-gas, or of fire-damp, the platinum becomes still hotter, and finally glows with a full red heat. Two spirals are made of this wire, one enclosed in a sealed tube full of air, the other fitted alongside of it, but exposed to any gaseous atmosphere into which the apparatus may be placed. On sending a mild electric current through both, they become heated, but the one in the "gassy" atmosphere glows much more brightly than the one in the sealed tube of air, and the greater the proportion of fire-damp present, the more brightly it glows. The appliance is very sensitive, and is said to detect  $\frac{1}{4}$  per cent. of gas.

A device depending upon similar principles (the action of hydrogen or fire-damp on platinum) forms the subject of a patent taken out by Mackie of Pennsylvania. The greater the proportion of fire-damp in the suspected atmosphere, the greater the distance from the lamp-flame, at which a coil of platinum wire begins to glow with heat. Through a sheath which passes through the base of the lamp, there slides a vertical graduated support, the top of which is bent so as to support the platinum spiral over the oil flame. When the lamp is brought into the suspected atmosphere, the support is pushed up or down, until the point is determined at which the platinum begins to glow,

and the corresponding percentage of fire-damp is shown by a graduated scale on the support.

The following test has been suggested also for the presence of fire-damp:—

If an ordinary tuning-fork, when sounded in air, gives out a certain note, it will be found that, when placed in a vessel containing carbon-dioxide, the sound is slightly louder. On the contrary, when placed in an atmosphere of hydrogen or fire-damp, the sound becomes fainter. In other words, the sound is affected by the density of the atmosphere in which the fork vibrates. Though perfect in principle, it is somewhat doubtful whether this test could be made of practical use in the mine.

**Purpose of the Gauze in a Safety Lamp.**—All metals are good conductors of heat, and, as we have seen, this property varies with different metals. (See under “Conduction.”) Now, the purpose of the gauze of a safety lamp—which is made of metal, and therefore is a good conductor—is to rob the gases of their heat. A certain temperature is necessary before any gas, or mixture of gases, will inflame or explode, and if the temperature be kept down below this point, then ignition cannot take place.

The principle upon which the safety lamp is based is well illustrated in Fig. 34. The gauze conducts the heat away from the flame so rapidly that the temperature does not rise to the point required to ignite the gas on the other side. The cooling effect is also helped by the number of small holes (mesh) of the gauze, which cause the heated gases to be spread over a large extent of cooling surface; this effect is greatly assisted by the action of Radiation (which see).

The gauze used in the safety lamp is usually of a standard size—namely, 28 meshes to the linear inch—and the wire forming the gauze is of 28 gauge in thickness, the metal used being iron or copper. This was the size of the gauze used by Davy (784 apertures per square inch). The

Belgian Mueseler lamp, authorised by the Belgian Government, has a gauze with 900 apertures per square inch. The original Marsaut lamp had a gauze with 934 apertures per square inch.

The old Davy lamp was not safe in an explosive mixture of fire-damp and air, if the current was moving at a rate of 5 or 6 feet per second. Modern lamps are protected by shields or bonnets, and are so arranged that it is impossible

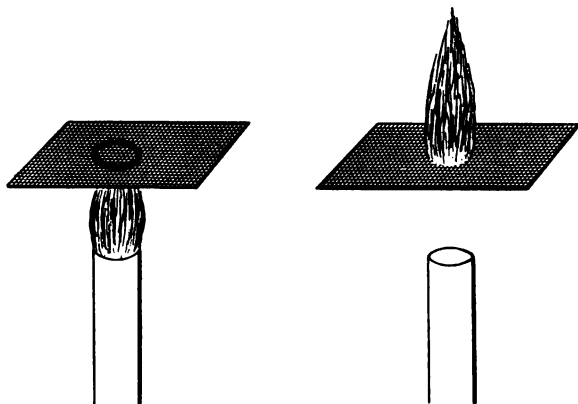


FIG. 34.—Illustrating the action of gauze on flame.

for a rapid current of air to pass through a lamp, even when exposed to a current of high speed. A properly constructed bonneted Mueseler should stand a current of 50 feet per second without passing the flame, and this represents about an average figure, though some lamps would withstand much more. The flame of a Davy lamp may not only be blown against the gauze so as to heat it considerably, and so make it unsafe, but the wave of air produced by a blown-out shot is quite capable of forcing the flame suddenly through the gauze, even at some distance from the place where the shot blew out; and if an explosive mixture should be present in the atmosphere at the time, disastrous results might follow.

**Ignition Point.**—It is necessary, before chemical action, resulting in combustion or explosion, can take place, that a certain temperature be reached, depending upon the nature of the chemical action, and upon that of the substances taking part in it. This temperature varies within very wide limits for the many chemical actions which are possible, and the point or temperature which must be reached in order to bring about these chemical changes is spoken of as the "ignition point."

Certain inflammable substances—for instance, turpentine, carbon-di-sulphide, ether, etc.—have low ignition points, that of carbon-di-sulphide being about 250° Fahrenheit; and, similarly, various gases and gaseous mixtures have their particular ignition point, below which temperature, combustion, or explosion cannot take place. The purpose of the gauze in the safety lamp is to help to keep the temperature of the gases, in an explosive mixture of fire-damp and air, below the ignition point, by conducting away the heat. The ignition point of an explosive mixture of coal-gas and air is 1200° Fahrenheit; that of fire-damp and air, 1230° Fahrenheit.

That combustion depends upon temperature may be proved by suddenly lowering the temperature of a burning gas, when it becomes extinguished. This principle is made use of in certain "safety-jet" appliances where the oxy-hydrogen flame is being used. In case of any back suction the gases are bound to pass through a space packed with metallic gauze, which absorbs their heat so rapidly as to bring them below their combustible or ignition point.

**Flash Point of Oils.**—Various illuminating oils—vegetable, animal, and mineral—are used in safety lamps. Some are of very low flashing points—that is, their vapour is inflammable at fairly low temperatures. For ordinary safety-lamps which do not burn spirit, but heavy oils such as colza, a mixture of two parts pure colza and one part best petroleum gives the best results. The

objection to the use of colza alone is the formation of a crust on the wick, after the lamp has burnt a while. The petroleum prevents this to a great extent, and, at the same time, a whiter light is given out. Two gallons of this mixture are sufficient to fill about 160 lamps of average size for 10 hours' burning.

There is naturally more danger in the use of an oil which has a low flash point than in one with a fairly high flash point, because the former gives off vapour readily, forming an explosive mixture with air. Unless under very special circumstances, none should be used having a flash point below 73° Fahrenheit. The flash point of mineral oils ranges from about 75° F. to 140° F. When low flash

oils are used, the practice is to absorb the oil in sponge or similar material, and not to fill the oil chamber with the liquid itself.

The flash point of oils is raised by exposure to the air. Lamps which burn paraffin or petroleum have been regarded with suspicion, and often justly so. There is also a tendency to compare the disastrous results often attending the use of an ordinary house lamp burning paraffin, with the possibility of similar results in the mine. Paraffin has got a bad name on this account, but there is really no comparison between the two cases.

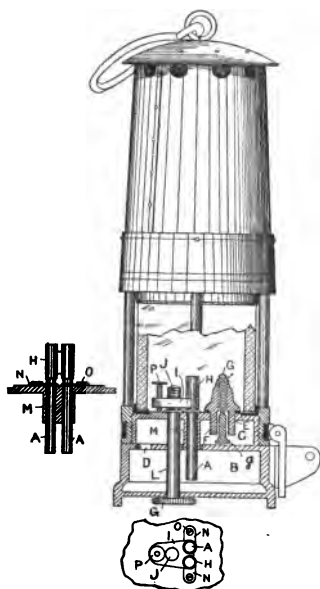


FIG. 35.—Naylor's safety-lamp.

A miner's lamp, burning paraffin only, has been devised by Mr Naylor, of Wigan. The oil chamber and wick holders are so constructed that the oil is quite unable to be spilled or upset, even if the lamp be held

upside down. It will be seen from the diagram (Fig. 35), that the tubes carrying the wick *aa* pass nearly to the bottom of the oil reservoir *B*, and the principle is really that of the "unspillable" ink-bottle. The burning of petroleum in this lamp is rendered safe by having a partition, or division plate, in the oil chamber, thus really converting it into two chambers. The wicks used are two in number and circular in form, and Fig. 36 shows a method of electrically igniting them, if desirable, but the lamp is also made without this attachment. The electrical ignition is effected by bringing the knob *κ* in contact with one term-

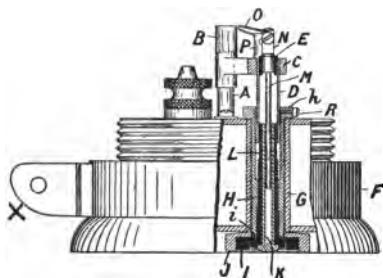


FIG. 36.—Electrical ignition of Naylor's lamp.

inal of a secondary battery or accumulator, and the part *x* in contact with the other terminal. This completes an electrical circuit, in which is a piece of platinum wire, *o*, held in position near the wicks. The passage of the current causes the platinum to glow with a full red heat, and ignites the wicks. The black portion, *i*, shows the insulation of the rod, *kn*, by means of vulcanite or ebonite. With a half-inch flame in this lamp,  $2\frac{1}{2}$  per cent. of fire-damp can be readily detected; with 11 per cent. the flame is extinguished. The lamp will burn 14 or 15 hours with practically no attention.

The candle-power from an ordinary size of miner's lamp varies from .25 to .75. The light emitted by this paraffin lamp is 1.1 candle-power, a very considerable advantage. Naylor's lamp has been described because of

its lighting power being much above the average, due to the use of petroleum combined with a special form of wick and wick holder. It is quite beyond the scope of this book to describe other lamps—such as the Marsaut, Mueseler, Hepplewhite-Gray, etc.—since they are dealt with in great detail in books dealing with coal mining.

**Lighting Powers of Lamps.**—The illuminating powers of different sources of light are compared by instruments called *photometers*, or *light measurers*. The standard, or practical unit for comparison, is a sperm candle (six to the pound) burning 120 grains per hour.

In Rumford's photometer, the lighting powers from two different sources are compared by adjusting the two lights



FIG. 37.—Rumford's photometer.

to be tested, so that they throw shadows of the same depth on a fixed screen. The shadow is cast by a vertical rod (see Fig. 37) which is placed between the screen and the sources of light. The brighter light will have to be placed at a greater distance from the screen, than the other. When the shadows are of equal depth, then the intensities of the lights will be directly as the squares of their distances from the screen. For

instance, suppose that one light is distant 12 inches and the other 18 inches from the screen, then the lighting powers are as  $18^2 : 12^2 = 324 : 144$ . Therefore the light, which is 18 inches distant from the screen, has  $\frac{324}{144} = 2.25$  times the illuminating power of that which is at 12 inches from the screen.

Most of the photometers in use at present are based upon a method originally used by Bunsen (Fig. 38). If a piece of white paper be soaked in melted spermaceti wax, so as to cover all the surface except a small circle in the centre, on holding it up to a source of light the portion

soaked with sperm is fairly transparent, and appears lighter than the unsoaked portion. If, however, the light be between the paper and the eye, then the soaked portion appears darker than the other, because the more opaque portion reflects the light better than the transparent part does. It follows, therefore, if the waxed paper be placed between two lights which are to be compared for intensity, that by moving the paper until it is equally

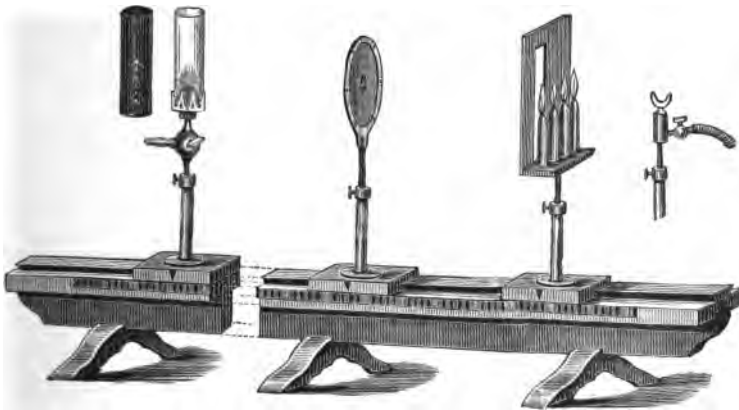


FIG. 38.—Bunsen's photometer.

lighted on both sides, no central spot will be visible. The distance of each light from the waxed paper is then accurately measured and squared, giving the comparative intensities of the two lights. The waxed paper is usually mounted in a box between two small mirrors placed at such an angle that the person making the test can see both sides of the paper at once, and thus adjust the distance to a nicety.

By means of these photometers, also, the candle-power of safety lamps can be gauged if desired.

The photometric tests of safety lamps are made by these means; but it must be remembered that they are made under the best conditions—namely, when the lamps are

perfectly clean : consequently the results thus obtained are always higher than can be expected when the lamps have been in use a short time, and exposed to the action of coal dust, grease, and dirt generally. Further, it is the horizontal rays from the lamp which are measured by a photometric test, whereas the most useful part of the light to the miner is that which is thrown upwards. In the Hepplewhite-Gray lamp a reflector is used, so as to throw the rays in the most useful direction.

Lamps burning acetylene gas (whose high illuminating power is well-known) are used on the Continent in mining, and are said to be more economical than electric lamps.

**Stink-Damp.** — Chemical formula  $H_2S$ . Molecular weight, 34. Density, 17. Specific gravity, 1.183. Composition by weight, 2 parts of hydrogen combined with 32 parts of sulphur. By volume—2 volumes hydrogen combined with 1 volume of sulphur vapour.

One of the earliest acquaintanceships which the chemical student makes in the laboratory is with the gas known there as sulphuretted hydrogen, and called by the miner *stink-damp*—a term denoting its most evident property, generally likened to the smell of rotten eggs.

It is a colourless gas, and has distinctly poisonous properties when inhaled in anything like a pure condition ; even when diluted with air it produces headache and giddiness. Compared with such gases as oxygen, hydrogen, fire-damp, white-damp, it is very soluble in water, the solution having similar properties to the gas. It is important to remember that it is a combustible gas, and in certain proportions may form an explosive mixture with air.

A familiar example of the action of this gas is seen in its effect upon metallic silver ornaments, which become covered by a bluish-black deposit after being exposed to the air for some time, due to the union of the silver with sulphur, forming silver sulphide. For similar reasons, paints consisting largely of white lead gradually turn

brown or black, owing to formation of lead sulphide. The decomposition of many organic substances containing sulphur results in the production of this gas along with others.

Though objectionable in many respects, stink-damp proves sometimes a good friend to the miner. Its characteristic smell is so readily noticed, even when in small quantities, that it cannot fail to attract attention. Now, the production of this gas is usually a forerunner of *gob-fires*, and is also an indicator of the coal heating spontaneously. Further, it is a valuable indication to the miner to be on the look-out for the deadly *white-damp*, which, together with *black-damp*, is always given off from *gob-fires*.

**Ventilation.**—From the consideration we have given to the properties of the various gases met with in mines, and the dangers they give rise to if allowed to accumulate, it will be readily understood how necessary is thorough ventilation to the safety and well-being of persons employed in the mine.

Not only is it necessary to have a sufficient atmospheric current to dilute and sweep out accumulations of fire-damp and noxious gases, but it is also essential that every human being and animal in the mine should have an ample supply of air for healthy breathing, and that the products of their breathing, and also of the burning of lamps, should be carried away. When doing no work, an average man breathes out at least 500 cubic inches per minute of respired air, which is rich in carbon di-oxide. Much more of the latter gas is contained in the breath if the man is working hard, because he is using up more carbon. Further, the surface of the coal should be kept cool by currents of air, so as to prevent spontaneous combustion.

It is evident, then, that a large volume of air will have to be introduced into the mine from without. The amount varies with conditions of the mine, whether deep or shallow,

hot or cold, gassy or otherwise, and may be anything from 100 to 500 cubic feet of air per man per minute. The powerful fans used for ventilating purposes at the present time in Great Britain are able to supply as much as 500,000 cubic feet of air per minute if necessary.

The subject of ventilation is a very important one, and in its entirety is quite beyond the scope of this book; at the same time its principles depend upon certain laws in physics, and therefore a brief outline may be attempted. The question of any kind of mine ventilation is a matter of effecting a difference, by some means, in the density of the air in two shafts. (The method of ventilation by furnace is almost obsolete now.) It depends upon heating the column of air in one shaft, and thus lowering its density and causing it to ascend. Charles's law explains the action in this case.

In ventilating by fan, the pressure is either lessened (by exhausting) or increased (by forcing) by mechanical means. In either case the action depends on Boyle's law.

In two shafts forming part of a ventilating system, the weight of the column of air in one exceeds the weight of the column of air in the other, by an amount which represents the "ventilating pressure" per square foot. This weight or pressure is that of the "motive column" as it is called, usually measured by an ordinary water gauge. The "motive column," then, is a column of air of such a height as will give a weight equal to the ventilating pressure.

Since 1 cubic foot of water weighs 62.4 lbs., and  $\frac{62.4}{12} = 5.2$  lbs. is the weight of 1 square foot of water 1 inch deep, then taking the weight of 1 cubic foot of air as .075 lbs.,  $\frac{5.2}{.075} = 69.3$  feet of air column equal to a weight of 5.2 lbs. as shown by 1 inch of water gauge.

This is approximately 70 feet of motive column per inch of water gauge.

The water gauge above mentioned is a glass tube, of U shape, often in metal fittings, so arranged that one of the limbs may be attached to a pipe connected with the fan-drift, or return air-way. Water is placed in the tube to a certain level, equal in both limbs, of course; by noting the effect on the level of this water produced by the pressure, the amount of the latter is measured, each inch denoting, as stated, 5.2 lbs. pressure per square foot.

**Combustion.**—In a wide sense, whenever chemical action goes on, accompanied by heat and light, combustion is said to take place, though in all cases the formation of heat and light is not necessarily brought about by chemical action. We usually speak of a substance being combustible *with regard to air*. For instance, we learn that ordinary coal gas is combustible, or will burn *in air*, but that coal gas will not support combustion nor allow a light to burn in it.

The very fact of the combustion of coal gas in air shows that there must be chemical action going on between them, and one naturally asks, If coal gas will burn in air, owing to chemical action, cannot the process be reversed and air be made to burn in coal gas? The answer to this question is, that it certainly can.

In Fig. 39, air is represented burning in an atmosphere of coal gas. The latter is supplied through the pipe *a* and is ignited at *b*, where there is an outlet in a small piece of asbestos mill-board placed on top of the lamp glass. Then on pushing a lighted taper up the tube *c*, through which air is being drawn, the latter takes fire and burns as shown.

To take a further illustration of the point, suppose that our atmosphere, instead of being composed of ordinary air as it is, were composed of coal gas (and that human existence were possible under such conditions), then if the gasometer's at the gas works supplied *air* to the houses instead of gas, on applying a light to the air issuing from

the pipes it would become ignited and burn, and the time-honoured custom of searching for an escape of *air* with a light would be attended by the same disastrous results as now—namely, explosion. It will be seen from this that the terms “combustible” and “supporter of combustion” are purely relative, and depend upon conditions existing at the time.

At a meeting of the Chemical Industry Society, in December 1894, the results of experiments by Clowes and Feillman were given, in connection with the power of certain gases to extinguish flames. They proved that carbon-di-oxide is a more powerful extinguisher of flame than nitrogen is, probably on account of its density.

Attention was called, further, to the great practical advantage which is derivable from the large proportion of carbon-di-oxide (58 per cent.) necessary to extinguish the hydrogen flame.

In the Clowes lamp (see Fig. 32), a hydrogen flame can be burnt at will, instead of the oil flame. It is found that by burning the hydrogen flame besides the oil flame, in parts of the mine where the atmosphere might contain large amounts of carbon-di-

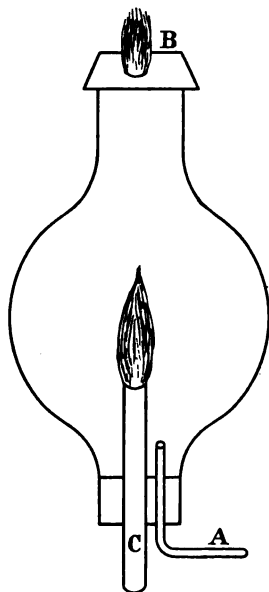


FIG. 39.—Air burning in coal-gas.

oxide, the extinction of the oil flame is not attended by loss of flame in the lamp, since the hydrogen flame never goes out, and serves to rekindle the wick in fresh air after the oil flame has been extinguished by a foul atmosphere. This has proved to be a great convenience to exploring parties in a mine after a fire or explosion, and is in conse-

quence of the power that the hydrogen flame has, to resist extinction in carbon-di-oxide.

**Spontaneous Combustion.**—The spontaneous combustion of coal is due to chemical action between the coal and the oxygen of the air. Other cases of spontaneous combustion occur, due to similar causes. Thus, heaps of oily cotton waste have been known to cause disastrous fires by firing spontaneously. The chemical action begins slowly at first, and develops a little heat; this heat is developed more quickly than it can be dispersed, and gradually increases, until it causes smouldering of the mass; and finally the heat rises to a point necessary for the chemical action to be accompanied by light and further heat. This may be illustrated more rapidly by means of phosphorus, a very inflammable substance. If a heap of small pieces be exposed to the air, fumes first come off (denoting chemical action), and while this is going on the heat is increasing, until the temperature is reached at which the phosphorus melts; more active combustion goes on; and finally the phosphorus bursts into flame and burns brilliantly.

A similar process goes on during the stages which finally bring about the spontaneous combustion of coal. It is a very remarkable fact, and one not easy of explanation, that the spontaneous combustion of coal goes on better in darkness than in light; other favourable causes are a warm and moist atmosphere. Vigorous ventilation tends to prevent the spontaneous combustion, because it cools the surface of the coal. It is sometimes stated that the presence of iron pyrites (coal brasses) in the coal, helps its spontaneous combustion, and to some extent this is no doubt true; but it must be remarked that in certain mines which are most troublesome in this respect there is a marked freedom from pyrites.

When the heating action commences it is usually indicated by stink-damp (sulphuretted hydrogen) being formed,

which is readily detected by its smell. This is, of course, produced by the sulphur in the pyrites, owing to the decomposition of the latter, in presence of moisture and warm air. For a similar reason, stink-damp is a forerunner of gob-fires, being in that case usually associated with the deadly white-damp (carbonic oxide).

Spontaneous combustion is usually only troublesome in the case of bituminous coals. Anthracite, and steam coals generally, are free from this liability.

**Coal Dust and its Action.**—Dr Philips Bedson, in a paper read before the British Association in 1893, pointed out that coal dust contains enclosed gases. He was specially led to investigate this property on account of observations which had been made, that the dust produced when screening coal at Ryhope Colliery, in county Durham, would often fire at the lamps used for lighting the screen, and thus produce a kind of explosion. It was found, on exposing the dusts to various temperatures in a vacuum, that gases were given off, varying in nature and amount with the temperature. At between 80° C. and 100° C., hydro-carbon compounds came off. In further investigating the same subject, Mr M'Connell had found that, even after the coal had been heated to 180° C. the dust from it, when crushed, gave off further amounts of hydro-carbon gases. Some dusts examined by him yielded carbon-di-oxide, oxygen, and nitrogen only; others contained a small amount of combustible gas in addition; while a third class gave off gas similar to that found with the Ryhope seam, and contained some other combustible gases (hydro-carbons) besides methane (fire-damp, commonly called).

No doubt the existence of these occluded or enclosed gases in coal dust plays an important part in determining its explosive action, since whereas the most explosive mixture of methane (fire-damp) and air is one of the former to ten of the latter, one volume of the hydro-carbons found

in this dust requires 23 volumes of air to form the most explosive mixture. Hence a more dilute mixture is more explosive.

Many explosions have occurred in which it was practically impossible to have had fire-damp as an explosive agent, and coal dust from certain seams appears to be almost as explosive as gunpowder. The richer the coal, and the freer from mineral matter (ash and dust), the greater the tendency to explosion. These conclusions were arrived at after a series of experiments made in a disused shaft at White Moss, Skelmersdale, by Mr Hall, H.M. Inspector of Mines. It was also concluded that a brisk ventilation has the effect of making these coal-dust explosions more probable, and more severe. Experiments were, at the same time, made with certain "high" explosives (see "Explosives"), and it was found that it was impossible to fire coal dust by their explosion in a dusty atmosphere, but with blasting powder (a "low" explosive) violent explosions occurred.

Now, as explained in the next chapter, gunpowder and blasting powder, containing a deficient supply of oxygen for complete burning, always generate carbonic oxide, and, as pointed out by Mr Orsman, the inference is, that carbonic oxide is the cause of explosions of coal dust under these conditions. In some cases more than 30 per cent. of the gases produced by blasting powder, when exploded, consists of carbonic oxide.

At the time of these experiments by Mr Hall in 1892, Mr Orsman devised an ingenious method of obtaining a sample of the gas immediately after the explosion.

The composition of the gas was found to be—

Oxygen . . . .	3·9 per cent.
Nitrogen . . . .	75·9 „
Carbon-di-oxide . .	12·1 „
Carbonic oxide . . .	8·1 „

Hence it was proved that the presence of carbonic oxide, in

addition to its highly poisonous action, has the effect of very much increasing the danger of an explosion of coal dust.

In Mr Hall's report to the Secretary of State on the results of experiments carried out in 1893, he states "that a very small quantity of coal dust in suspension in the air would, along with that at rest on the timbers, be sufficient to bring about conditions, under which a colliery explosion might happen, through the firing of a gunpowder shot; and even without dust in suspension, enough could be dislodged from timbering, and other lodging places, to cause violent explosions. At the same time, he does not imply that all colliery explosions are due to dust, many being undoubtedly due to fire-damp, but these are never so destructive nor so widespread as the ones caused by coal dust."

We gather from this that an explosion may occur at one particular point, owing to the presence of fire-damp, and the explosive wave so formed may ignite, in the immediate vicinity, a mixture of air and coal dust, and so be carried on indefinitely, so long as air and coal dust are present. It has been found that the primary explosion causes a disturbance in the air, similar to that of a sound wave, and this wave, travelling in front of the explosive mixture, causes fresh dust to come into suspension in the air; consequently this explosive wave feeds itself, so to speak, from the disturbance of the atmosphere which goes on in front.

It has also been noticed that coal dust renders an explosive mixture of fire-damp and air more sensitive to explosion, and even one per cent. of fire-damp can be caused to burn when fine coal dust is present. This would point to the necessity of having a lamp which will indicate the presence of such an amount of fire-damp, since an ordinary lamp will only recognise as little as 2 or  $2\frac{1}{2}$  per cent. (see Clowes' lamp).

And it may be pointed out here, as stated under "Explosives," that various explosives—for example, those of the gunpowder and nitro-glycerine class which are

deficient in oxygen—are capable of producing gases containing a large proportion of carbonic oxide, which have the same effect as fire-damp in promoting explosive mixtures with air, especially in the presence of small quantities of coal-dust.

## CHAPTER IX.

### EXPLOSIVES.

**Nature of Explosives.**—Explosives are chemical compounds which are able to be very rapidly decomposed, upon the application of heat or shock. They suddenly generate, when thus decomposed, large volumes of gas, which at the same time are heated to a high degree, and are thus caused to expand greatly (see under Boyle's law). Most explosives in practical use are solid, a notable exception being nitro-glycerine.

The necessary conditions for an explosive mixture are to have some substance, or substances, which will readily burn under suitable conditions, mixed with some other substance which will supply the necessary material for the burning or combustion—namely, oxygen. An explosive must therefore contain all the necessary ingredients for a rapid combustion.

Many substances rich in carbon have been suggested or used as explosives, when mixed with an oxidising agent. For instance, charcoal, sawdust, bran, resin, starch, sugar, wood pulp. Coal-tar products, etc., have been used along with oxidising agents, such as nitrates, chlorates, picrates, permanganates, etc. The science of explosives has been developed upon these broad principles, and the result is that a large variety of mixtures, each suited for its own particular work, are available for practical use.

Many explosives do not contain enough oxygen to completely burn up the substances contained in them—for instance, gunpowder, gun-cotton, and nitro-glycerine. The result is the production of carbonic oxide (white-

damp). This is very undesirable, not only on account of its very poisonous action (see under Carbonic oxide), but also because the presence of this gas in an atmosphere very much increases the danger of an explosion of coal-dust. The reason for keeping down the supply of oxidising agents in the explosive, is to prevent too much heat being produced, and in attempting to tone down this temperature certain substances—such as wood-meal—have been added; this also causes the production of carbonic oxide. Nitrated oils—for example, linseed and castor oils treated with nitric acid—have been suggested as suitable for use in modifying the temperature, and reducing the shattering effect, produced by certain explosives.

When ammonium nitrate is employed as an oxidising agent it may easily be used in excess—that is, more than enough to supply oxygen to the materials to be oxidised—because on being decomposed it *absorbs* heat.

An explosive intended for use in coal or other fiery mines should leave no solid matter after explosion—that is, the products should all be in the gaseous condition; and, further, these gases should be harmless, free from noxious properties of all kinds. It should also, in actual use, when properly tamped, be unable to fire an inflammable or explosive mixture of fire-damp and air, or of coal-dust and air.

It is generally admitted that the use of “high” explosives lessens the risk of explosion in mines. The mechanical effect of the explosion is also an important point; if it be too shattering, then the coal is broken up too much, which is an evident disadvantage; a tearing or rending effect is required, so as to break down the coal in sizes most suitable for commercial purposes.

Explosives may be classified thus:—

- (a) Those depending upon the use of nitro-glycerine, either mixed with some absorbent substance, or gelatinised by adding varying amounts of nitro-cellulose.

- (b) Those in which gun-cotton or nitro-cellulose is the basis.
- (c) Those in which a hydrocarbon, preferably nitrated, is used, along with an agent rich in oxygen, such as ammonium nitrate, often known as "Sprengel explosives."

All "high" explosives require a detonator to explode them, and the explosive effect is brought about by the combined action of heat and shock. "Low" explosives can be exploded with a fuse, without the use of a detonator. Some explosives require more powerful detonators than others. A detonator is a small metal cap, generally of copper, containing a substance called fulminate of mercury, usually mixed with a little chlorate of potassium. The fulminate is prepared by dissolving mercury (commonly called *quicksilver*) in nitric acid. The solution is cooled down and alcohol added to it, this causing fulminate of mercury to be produced as a powder in the liquid. The solid is separated by suitable means, and thoroughly washed to purify it, and free it especially from acid. It is then carefully dried and highly compressed, forming a material which is most sensitive as regards explosion. The slightest friction, or moderate heat (about 300° F.), is sufficient to explode it.

Detonators can be fired with a time fuse *if the conditions of the mine permit*, but it must be remembered that in using a fuse it may give out sparks which are sufficient to ignite an explosive mixture of fire-damp and air. Further, it is obviously absurd to use a fuse which may bring this about, when using a *safety* explosive.

Undoubtedly the best and safest means of firing a detonator is the use of electricity. This may be done by a *high-tension* or *low-tension* current. In the former case a small current, at a high pressure, is used, by means of which electric sparks are caused to leap across a gap between the terminals of two wires, and thus ignite the "priming" with which they are surrounded. This in turn

causes the fulminate to explode, and to fire the cartridge. The priming is itself an explosive, depending, as all others do, on a substance to be oxidised (usually in these cases, sulphide of antimony) and an oxidising agent, usually potassium chlorate. Fig. 40 shows one form of high-tension exploder. In the case of low-tension firing, a comparatively large current at a low pressure is used, and this is made to pass through a fine platinum wire, which is thus raised to a red heat, owing to its

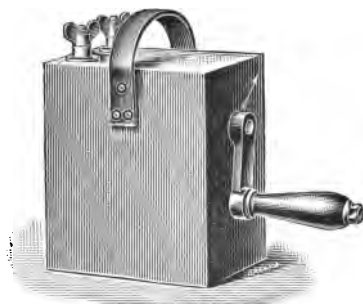


FIG. 40.—High-tension Exploder.

resistance to the passage of the current; this fires the priming and explodes the detonator.

The composition, properties, and preparation of some of the most commonly used explosives may now be considered.

**Gunpowder**, until comparatively recent years, had an average composition of potassium nitrate, 75 parts; charcoal, 15 parts; and sulphur, 10 parts; but this composition is now often much modified, and varies considerably (see Bobbinite and Special Bull-dog).

There are many objections to the use of gunpowder for mining purposes. Since it does not contain enough oxygen to burn the charcoal and sulphur completely, there is always carbonic oxide (white-damp) produced by its explosion (see

under Carbonic oxide for conditions under which this gas is produced). The composition of the gases resulting from an explosion of gunpowder is:—

Carbon-di-oxide . . .	51.3 per cent.
Carbonic oxide . . .	3.5 „
Marsh gas and hydrogen . .	3.5 „
Nitrogen . . . . .	41.7 „

Further, it produces much flame by its explosion, a most undesirable property in a flery or gassy mine. As a compensation for these drawbacks, it has the advantage in a coal mine of breaking down the coal in large masses, owing to the comparative slowness with which it explodes. According to Trauzl, 2.2 lbs. of gunpowder, occupying a space equal to a cube of 4 in. side, develops  $1\frac{1}{2}$  million foot lbs. in  $\frac{1}{100}$  part of a second; the same quantity of dynamite, occupying a space equal to a cube of  $3\frac{1}{2}$  in. side, develops over  $7\frac{1}{4}$  million foot pounds in  $\frac{1}{50,000}$  part of a second. As a result of this, gunpowder fired on a stone slab, say, does not injure the slab, since the slow combustion enables it to spend its force in the air; but dynamite explodes so quickly that it has not time to expend its force thus, but shatters whatever it may be in contact with.

**Gun-cotton** is, chemically, a nitro-compound of cellulose, made by the action of nitro-sulphuric acids on cotton fibre. Unless very special precautions are taken in its manufacture, the cotton fibres retain small quantities of the acids used, and the material is then liable to explode spontaneously. Many disastrous accidents occurred before this was recognised. The actual oxidising agent is the nitric acid, the sulphuric acid being used simply to take up the water which is produced during the chemical reaction, in oxidising the cellulose.

Gun-cotton is a compound containing too little oxygen for complete combustion, and therefore carbonic oxide is

produced by its explosion. It is necessary to use a detonator to bring about explosion, which is accompanied by much heat, causing enormous expansion of the gaseous products. It is never used in the mine alone as a blasting agent, although mixed with barium nitrate it constitutes the explosive Tonite (which see).

**Nitro-glycerine** is made by the action of nitro-sulphuric acid on glycerine. It is a liquid, and in this form there are many objections to its use; it is not used for blasting purposes in the mine. Nobel used a solid substance called *kieselguhr* to absorb or soak up the liquid. *Kieselguhr* is a siliceous earth found on the Continent, and is principally composed of the shells of minute sea animals. The mass formed by the mixture of this earth with nitro-glycerine is the well-known explosive dynamite.

In some cases, more active substances are used instead of *kieselguhr* to mix with the nitro-glycerine, such as sulphur, magnesia, carbon in some form (sawdust, wood-pulp, etc.), also nitrates and chlorates. Such mixtures constitute very powerful explosives, known under a variety of names (Rack-a-rock, Hellhoffite, etc.) which are principally used for blasting or shattering purposes.

**Gelignite** consists of a mixture of nitro-glycerine, nitro-cellulose, wood-meal, and nitre.

**Tonite** has for its basis gun-cotton, with which is mixed nitrate of barium, the use of the latter being to supply oxygen to the mixture, since the gun-cotton itself does not contain sufficient for complete combustion.

**Roburite** contains 86 to 89 per cent. of ammonium nitrate, mixed with 9 to 13 per cent. of di-nitro-benzol, and small amounts of chloro-naphthalene (see List of Permitted Explosives, pp. 94-96).

The gases produced by its explosion consist of 32 per

cent. carbon-dioxide, and 68 per cent. nitrogen, and are therefore free from the objectionable and highly dangerous carbonic oxide.

**Ammonite** consists of  $87\frac{1}{2}$  per cent. ammonium nitrate and  $12\frac{1}{2}$  per cent. di-nitro-naphthalene, another organic coal-tar product.

The gaseous products of the explosion of this compound are free from carbonic oxide, and consist of carbon-dioxide and nitrogen.

**Bellite** is yet another admixture containing ammonium nitrate, and an organic body—di-nitro-benzol. Its gases are free from carbonic oxide, and are entirely carbon-dioxide and nitrogen.

In **Westphalite** the substance to be oxidised is resin, present to the extent of about 5 per cent., mixed with 95 per cent. ammonium nitrate. The gaseous products are the same as from the three last-named explosives.

The use of ammonium nitrate is one of the comparatively recent developments in the manufacture of explosives. It is especially useful in several ways. There is no solid residue left after heating or exploding; when decomposed it does not give out heat, but absorbs it, and therefore sufficient can be used in a mixture to completely burn up or oxidise the materials used, without producing an excessive amount of heat in doing so. This is the reason why no carbonic oxide is produced, by an explosion of those compounds containing this nitrate. Again, the ammonium nitrate explosives are not so quick in their action as the nitro-glycerine compounds are, and are therefore more suited for breaking down coal.

The principle upon which the manufacture of smokeless powders is based is to have a substance, such as nitro-cellulose, which leaves no ash when burnt (that is, its products are entirely gaseous), and a nitrate, such as ammonium, which also leaves no solid residue. Hence

such substances can have nothing but gaseous products when exploded, and these gases, containing no solid matter in suspension, are invisible.

**Lime Cartridges.**—In very fiery or gassy mines these are sometimes used. They consist of quicklime (that is, freshly burnt lime) compressed into cartridge form; they are usually about 5 in.  $\times$   $2\frac{1}{2}$  in., and have to be kept in air-tight cases (as the limes used for oxy-hydrogen lighting have); otherwise the moisture and carbon-di-oxide which are absorbed from the air rapidly spoil them for their purpose. Their action depends upon the “hydration” of the quicklime—that is, its combination with water. The oxide of lime becomes converted into hydrate of lime, and in doing so expands with great force, and breaks down the coal. No explosive effect or flame is produced. Objection is sometimes made to the appearance of the coal brought down by these means, but its quality is not affected at all. Though absolutely safe, the use of lime is nothing like so extensive as that of the “safety” explosives.

One of the suggested uses to which liquid air could be put, is as an explosive agent. In combination with carbonaceous substances it forms an explosive mixture. An objection to its use would be that the cartridge would have to be fired in a short time after removal from the liquid air. It has been used, but so far results have not been very encouraging.

As we have already seen under “Water,” when a current of electricity is passed through water, the latter is decomposed into the gases oxygen and hydrogen; these are spoken of as electrolytic gases, and have been used as an explosive agent on the Continent. The force generated when these mixed gases are exploded is very considerable, but they have the disadvantage, in some cases, of causing the “firing” of a gassy or explosive atmosphere. The actual cost of the gases is not great.

**Permitted Explosives.**—The following is a list of some of the explosives officially sanctioned for use in coal mines, which will serve to illustrate the different types:—

*Ammonite*—

Nitrate of Ammonium	.	87 to 89 per cent.
Di-nitro-naphthalene	.	11 „ 13 „
Moisture	.	— „ .5 „

*Amvis*—

Nitrate of Ammonium	.	88 to 91 per cent.
Wood-meal	.	4 „ 6 „
Moisture	.	— „ .5 „
Di-nitro-benzol	}	4 „ 6 „
Chlorinated Naphthalene		

*Carbonite*—

Nitro-glycerine	.	25 to 27 per cent.
Nitrate of Barium	}	30 „ 36 „
Nitrate of Potassium		
Wood-meal	.	39 „ 42 „
Sulphuretted Benzol	.	— „ .5 „
Carbonate of Sodium	}	— „ .5 „
Carbonate of Calcium		

*Dragonite*—

Nitro-glycerine	.	34 to 37 per cent.
Nitro-cotton	.	2 „ 3 „
Nitrate of Potassium	.	43 „ 46 „
Vaseline	.	5 „ 6 „
Wood-meal	}	11 „ 13.5 „
Charcoal		

*Nobel Ardeer Powder*—

Nitro-glycerine	.	31 to 34 per cent.
Kieselguhr	.	11 „ 14 „
Sulphate of Magnesium	.	47 „ 51 „
Nitrate of Potassium	.	4 „ 6 „
Carbonate of Ammonium	.	— „ .5 „
Carbonate of Calcium	.	— „ .5 „

*Bobbinite—*

Potassium Nitrate . . .	62 to 65 per cent.
Charcoal . . . . .	17 „ 19.5 „
Sulphur . . . . .	1.5 „ 2.5 „
Sulphate of Ammonium } . . .	13 „ 17 „
Sulphate of Copper . }	
Moisture . . . . .	— „ 2.5 „

The sulphate of ammonium not to exceed 11 per cent., by weight, of the finished explosive.

*Coronite or Permitite—*

Nitro-glycerine . . . . .	38 to 40 per cent.
Nitro-cotton . . . . .	1 „ 1.5 „
Nitrate of Ammonium . . . . .	26 to 28 „
Nitrate of Potassium . . . . .	3 „ 5 „
Stearate of Aluminium . . . . .	11 „ 14 „
Rye-flour . . . . .	8 „ 11 „
Wood-meal . . . . .	2 „ 4 „
Liquid Hydrocarbon of Paraffin } series . . . . .	2 „ 4 „
Moisture . . . . .	— „ 2.5 „

The hydrocarbon is to have a flash-point of not less than 200° F.

The stearate of aluminium is to be free from mineral acid.

*Westphalite—*

Nitrate of Ammonium . . . . .	94 to 96 per cent.
Resin . . . . .	4 „ 6 „
Moisture . . . . .	— „ .5 „

*Negro Powder—*

Nitrate of Ammonium . . . . .	86 to 90 per cent.
Tri-nitro-toluol . . . . .	9 „ 11.0 „
Graphite . . . . .	1 „ 3 „
Colouring matter . . . . .	— „ .1 „
Moisture . . . . .	— „ 1.0 „

*Roburite No. 3—*

Nitrate of Ammonium . . . . .	86 to 89 per cent.
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Di-nitro-benzol . . .	9 „ 13 per cent.
Chloro-naphthalene . . .	— „ 2 „
Moisture . . .	— „ .5 „

The chloro-naphthalene to contain not more than one part of chlorine.

*Special Bull-dog—*

Nitrate of Potassium . . .	84 to 86 per cent.
Carbonate of Magnesium . . .	2.5 „ 3.5 „
Charcoal . . .	12 „ 13 „
Moisture . . .	— „ 2 „

*Ammonal—*

Nitrate of Ammonium . . .	94 to 96 per cent.
Metallic Aluminium . . .	2.5 „ 3.5 „
Charcoal . . .	2 „ 3 „
Moisture . . .	— „ 1 „

**Apparatus for use in Atmospheres which will not support Life.**—The necessity of having to go into the workings of a mine, after an explosion of fire-damp, to render aid to those who have suffered from the direct effects of the explosion, or have been overcome by the gases resulting from it, has led to the device of appliances called pneumatophores, by using which a person is able to exist in any atmosphere of noxious gases, for a reasonable time. These consist, in principle, of a kind of mask or head-dress similar to that which a diver uses. By means of tubes and valves suitably arranged, the wearer is able to supply himself with oxygen, which he carries in a compressed form in a strong steel, or copper, cylinder. At the same time the carbon-di-oxide which he is continually breathing out, is being absorbed by caustic potash, in the form of sticks or rods, and contained in a bag connected with the mask. The rods of potash form a large absorbing surface, and take up the carbon-di-oxide almost as quickly as it is breathed out. The nitrogen in the original air remains constant and unaltered throughout, and serves its usual

purpose of diluting the oxygen, supplied from the gas cylinder.

In an Austrian colliery much subject to gob fires, eighty sets of apparatus of this kind were in constant use.

The oxygen gas contained in the cylinders is under a pressure of one hundred atmospheres. It is necessary to use the apparatus with caution, and only experienced men should make use of them, otherwise they may prove veritable death traps. Their limitations should be known, and the wearer should always leave himself an ample supply of oxygen for his return journey from any exploring. It is in fact advisable that two oxygen cylinders be used, one for the outward journey, and the other for the return. It has been suggested, too, that compressed air be used instead of oxygen, and that it be discharged from the apparatus instead of breathing it over and over again.

A weak point in connection with the use of compressed oxygen is that it often contains nitrogen, in which case the latter gradually accumulates, and may cause the wearer to be overcome before he is aware of the danger.

Pneumatophores are very useful appliances, but it should be realised that they have their dangers.

## CHAPTER X.

### COMPOSITION OF VARIOUS COALS AND FUELS.

**Nature of Coal.**—All the different kinds of coal have been produced by the decomposition of vegetable matter which grew and flourished in remote ages. The chemical changes have occupied very long periods, and have been complex, the final products of these changes being coal in its various forms, water, marsh gas (fire-damp), and carbon-di-oxide. These pent-up gases are continually being met with in the coal measures and adjoining strata.

All coals contain varying quantities of the following substances: moisture, ash or mineral matter, sulphur, carbon, hydrogen, oxygen, and nitrogen—some of these being valuable constituents, others the opposite. To form an estimate of the value of a particular sample of coal, it is necessary to know the proportions of moisture, ash, sulphur, volatile hydrocarbons, fixed carbon, its calorific value (or heating power), and also its specific gravity—that is, its weight compared with water, bulk for bulk—from which an estimate may be formed of the space which a certain weight of coal will occupy. The determination of these constituents is described later.

What is called an ultimate or elementary analysis of coal would include carbon (total), hydrogen, oxygen, nitrogen, sulphur, ash, and moisture. Though the knowledge of total carbon is for some purposes useful, most generally a knowledge of the amounts of fixed carbon and volatile hydrocarbons is more important, because one coal may contain a greater proportion of carbon than another, and still yield less coke.

It is now proposed to describe the distinguishing properties of different kinds of coals.

**Anthracite** contains from 90 to 95 per cent. of total carbon, which is a greater amount than in any other kind of coal. It is probably bituminous coal which has been considerably altered by the action of heat and pressure. A seam containing anthracite may merge into one of bituminous coal at some point, just as cannel and bituminous coals are known to exist in the same seam. It does not soil the fingers like ordinary coal does, and gives out very little flame or smoke on burning. On heating, it decrepitates or crackles. Anthracite is hard and dense, breaking with a hollow, shell-like fracture. Its chief use is as a steam coal, though it is also used as a source of carbon, in re-carburising steel during its manufacture.

Anthracite coal is not so dangerous to work in the mine as bituminous coal, since the physical condition of the former does not allow pent-up gases to escape as freely as they do from the latter class.

Analyses of anthracite coal follow:—

	I.	2.
Carbon . . .	89·96 per cent.	88·72 per cent.
Hydrogen . . .	3·10 „	3·41 „
Oxygen . . .	1·63 „	2·54 „
Nitrogen . . .	·27 „	·39 „
Sulphur . . .	1·06 „	·87 „
Ash . . .	1·96 „	1·56 „
Moisture . . .	2·00 „	2·50 „

As a rule, anthracites contain from—

88 to 95	per cent.	Carbon
2 to 4·0	„	Hydrogen
1·5 to 3·0	„	Oxygen
·25 to ·75	„	Nitrogen
·5 to 1·5	„	Sulphur
1 to 3	„	Ash

An analysis of anthracite also gave the following results—

Fixed carbon	. . .	84·60 per cent.
Volatile hydrocarbons	. . .	12·40 „
Sulphur	. . .	·95 „
Ash	. . .	2·00 „

**Cannel** has a dull black appearance, not so bright as anthracite or bituminous coals. Its chief use is in gas manufacture, because of its richness in hydrogen. Cannel contains from 40 to 50 per cent. of volatile hydrocarbons, and one ton of Wigan cannel produces about 10,850 cubic feet of gas, other cannels ranging from 9000 to 15,000 cubic feet; the lighting power of the gas varies from 20 to 40 candle power. It is remarkable that the analysis of many cannels gives a similar composition to that of other coals; the difference appears to be largely a physical one, probably due to the combined action of heat and pressure.

Analyses of cannel gave the following results:—

	1.	2.
Ash . . . . .	2·30 per cent.	1·60 per cent.
Sulphur . . . . .	1·30 „	·58 „
Volatile hydrocarbons	40·56 „	48·85 „
Fixed carbon . . . . .	56·50 „	49·25 „

An ultimate analysis of cannel would average as follows:—

Carbon . . . . .	80·0	per cent.
Hydrogen . . . . .	5·5	„
Oxygen . . . . .	8·0	„
Nitrogen . . . . .	2·0	„
Sulphur . . . . .	·75 to 1·5	„
Ash . . . . .	2 „ 4	„

The ash in some cannels is very high, reaching as much as 20 to 25 per cent. The coke resulting from cannel is usually much inferior to that from a “caking” coal.

**The Bituminous or Flaming Coals** are divided into a *caking* and *non-caking* class. A caking coal is one, as the name implies, the particles of which tend to soften and fuse together into one mass on heating. As a rule non-caking coals, or such as do not alter their shape on heating, contain more carbon and hydrogen than the caking varieties. A caking coal produces much better coke than a non-caking, or than those which burn freely.

Certain coals seem to combine some of the properties of both caking and non-caking fuels, and the kind of coke left by them varies considerably in physical properties, such as density, compactness and friability. Non-caking coals are sometimes called "splint" coals; they splinter up on heating, and do not conglomerate or mass together. The caking appears to be a physical peculiarity of the coal, since with coals having almost the same chemical composition, some will cake and others will not.

For steam and house coals, the desirable qualities are the production of as much heat as possible with little residue of ash. Steam coals are such as are mostly used for the generation of steam, and generally burn freely, with comparatively little smoke; a good house coal should be readily set alight, and be fairly free from ash. In both kinds of coal, sulphur compounds should not be high, and the smoke produced by their combustion should not be too great.

With regard to the latter point—smoke production—this is to a great extent controllable by the consumer, especially in steam-raising. As a general rule, the higher the amount of volatile hydrocarbons, the greater the tendency to produce smoke, *if the conditions of firing are the same*. The anthracites, for instance, burn with practically no smoke, and have a correspondingly low volatile matter. But with coals, generally speaking, everything depends upon the firing. Careless firing, with a coal of low volatile matter, will produce more smoke than a high volatile coal more carefully and economically stoked.

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of silvery appearance, largely composed of carbon, and practically free from ash.

All the mineral matter present in the original coal remains in the coke. Its proportion is of course increased, since volatile matters have been driven off during the coking. For instance, if a coal contains 5 per cent. of ash, and gives off on coking, 50 per cent. of volatile matter, it is evident that the amount of ash in the coke will be 10 per cent., or double that in the original coal.

The sulphur of the coal is not all retained by the coke, some passing off in a volatile form on heating, and being included with the volatile hydrocarbons. There is no fixed rule as to how much is retained by the coke, as this varies with the kind of coal used. The only means is to test the coal for sulphur, and to test also the coke resulting from it. The sulphur driven off in coking is mainly the organic sulphur, together with some derived from the pyrites in the coal.

The phosphorus in coke is rarely present in very appreciable quantity—from traces to .02 per cent.—though in some districts the coal may yield a coke containing .10 to .15 per cent. Such an amount would be a very serious objection to the use of the coke as a metallurgical fuel, especially in iron smelting.

Coke always retains small amounts of hydrogen, nitrogen, and oxygen, so that all the nitrogen present in the coal is not obtained in recovery processes for conversion into ammonia sulphate or chloride. The amount found in the coke ranges from about .3 to 1.0 per cent.

An analysis of an average quality of coke is as follows:—

Ash	.	.	.	.	.	6 to 12 per cent.
Sulphur	.	.	.	.	1 „	2.5 „
Volatile matter	.	.	.	.	nil „	2.0 „

**Peat** is an accumulation of partially decomposed

vegetable matter, a damp, marshy ground and moderate temperature being favourable conditions for its production. In its ordinary condition it retains a large amount of moisture (at least 20 to 25 per cent.), but its general composition varies with its age, an older formation containing more carbon than a more recent one does. Peat also contains a large amount of ash or mineral matter, in which are found phosphates and sulphur compounds, so that it cannot be used in the manufacture of high grade iron, even if the objectionable moisture be not considered, and if the physical condition of the peat itself be altered so as to make it less friable and tender.

**Lignite** is a kind of intermediate product between peat and bituminous coal. It is darker in colour than peat, and is richer in carbon; it retains much moisture, and is worked more on the Continent than in the British Isles. In the change from vegetable matter to bituminous coal, the carbon gradually increases from about 50 per cent. in wood to 75 to 85 per cent. in bituminous coals, and 95 per cent. in anthracite.

**Wood.**—The chief constituent of wood is a substance named *cellulose*, which consists of carbon, hydrogen, and oxygen. The amount of mineral matter or ash it contains varies with the kind of wood, but is rarely more than 3 or 4 per cent., often as low as 1 per cent. It retains a good amount of moisture, but is not so objectionable in this respect as peat is, being also more readily dried.

The carbon it contains is best utilised in the form of charcoal, which is made by heating the wood either out of contact with air, or with admission of as little air as possible. It is very useful in metallurgical operations, where a pure fuel is desirable, and where substances like sulphur and phosphorus would completely spoil the product.

**Liquid Fuels** have an advantage over coal in districts where oil is plentiful. The oils are all *hydrocarbon* compounds, and most of them, petroleum oils in particular, have a greater heating effect than anthracite.

They burn when supplied with oxygen, forming carbon-di-oxide and water. They have the very great advantage of being practically free from ash, and contain very little sulphur (almost nil), though in some cases this may be as high as 4 or 5 per cent., and, moreover, being usually sprayed into the combustion chamber, a very complete combustion is possible. Very little, if any, of the oil escapes burning, and consequently the greatest possible heat efficiency is obtained.

#### **Undesirable Elements in Coal and other Fuels.—**

There are certain elements and compounds the presence of which affects the value of coal for many purposes.

If oxygen be present in a fuel, the heating effect of the latter is lowered, since, being combined with the carbon and hydrogen, these substances do not give out their full amount of heat. Had they not been already combined with oxygen they would have been at liberty to unite with the oxygen *of the air* and thus produce heat.

Some coals are practically free from oxygen, others contain 10 per cent. or more.

**Ash**, if present in appreciable quantity, is very objectionable in a coal, because it tends to form "clinker," a fusible compound, produced either by the ash itself, or by the combination of the substances composing the ash, with oxide of iron from the firebars. This clinker prevents free passage of air through the fuel, and tends to the production of carbonic oxide, and therefore loss of heat. It is evident, too, that the higher the amount of ash, the lower the calorific or heating value of the coal.

**Sulphur and Phosphorus** are both objectionable,

especially if the fuel is to be used for metallurgical work, since practically all the phosphorus and a large proportion of the sulphur go into the resulting metal and contaminate it. Sulphur is further objectionable, because it forms corrosive gases, and tends to become converted into sulphuric acid. The escaping gases from a sulphury fuel, also, contaminate the atmosphere and act injuriously on animal and plant life in the neighbourhood.

Sulphur exists in three different conditions in coal:—

(1) The most common form is as iron pyrites (chemically bisulphide of iron  $\text{FeS}_2$ ), commonly called “coal brasses,” and is, as the name implies, the “brassy” looking mineral often seen in coal.

(2) As sulphate of lime, the common name for which is “gypsum,” also sometimes as sulphate of alumina.

(3) As an organic compound with carbon and hydrogen.

Of these forms, the sulphur existing as sulphate of lime is the least objectionable. It is “fixed” by the lime, so to speak, and does not produce any noxious or corrosive gases on firing.

The iron pyrites is objectionable, not only because of the corrosive and noxious gases it produces, but because the remaining iron (in the form of oxide) fuses with the silica of the coal ash and increases the amount of clinker and helps to make it more readily fusible. Devices for getting rid of this objectionable element form the basis of many patents. The most successful methods depend upon washing the coal and separating the shaley and pyritic portions by means of their specific gravity.

The sulphur, which is in organic combination, is mostly converted into bisulphide of carbon through coming in contact with heated carbon during combustion. This refers principally to processes of coking and of gas manufacture where destructive distillation is going on, with little or no air present.

**Nitrogen in Coal.**—The element nitrogen cannot

be classed as an undesirable one in coal. The amount varies from .25 to 3.0 per cent. as a rule. It is probably due to animal remains from the periods at which the coal measures were in process of formation, and also some of it is derived from the atmosphere existing at the time. When the coal is heated, this nitrogen combines with some of the hydrogen of the fuel and forms ammonia (chemical formula  $\text{NH}_3$ ). It will be readily seen that in making coke from coal, considerable quantities of this compound, ammonia, must be produced.

Of late years, the recovery of this important compound (formerly allowed to escape and waste) has become a very important problem in the economy of a coking plant, and very few modern works can afford to neglect the recovery of this and other by-products, such as tar and various oils. The ammonia is first given from the coal, on burning, as a gas. This is converted into a solid by passing it through a solution of sulphuric acid or hydrochloric acid, whereby sulphate or chloride of ammonia is produced in a marketable form. These compounds are then used for the manufacture of ammonia salts, and also largely for manurial purposes, the nitrogen which they contain forming a valuable plant food.

**Briquettes.**—The manufacture of briquettes is a useful process, especially in working coal seams of inferior quality, where the waste may sometimes amount to 20 per cent., and also for converting into a workable form small coal of all kinds which is unavoidably produced in the mining and subsequent handling of coal.

In briquette form such coals are much improved, both for use and storage and also for transport. The space occupied by them is about one-half that required for coal of the same weight.

Some kind of binding material has to be used along with the fine coal. The most generally successful agent has been coal-tar, or pitch, or both. The subject forms the

basis of many patents, and all manner of binding materials have been suggested, such as glucose, waste glue products, starchy compounds, molasses, silicate of soda, oily materials, waste oil, crude linseed oil, sea-weed extracts, and gelatinous substances of all kinds. Resin is often successfully used alone or with other binding material.

It is preferable that the binding material used should, if possible, help or increase the burning qualities of the briquette, and should itself be as free from ash as possible. Magnesia cement is said to make a very strong briquette, but it has the disadvantage of increasing the amount of ash, though this increase need not be more than about 3 per cent.

Briquettes can be used as fuel for locomotives as well as for domestic purposes.

## CHAPTER XI.

### METHODS OF ANALYSIS OF COAL.

**Specific Gravity.**—By the specific gravity of a solid is meant its weight compared with an equal volume of water at a standard temperature.

The specific gravity of coal may be determined by first weighing a piece in air on a chemical balance, and then in water, suspending it in the latter by means of a looped horse hair.

The coal loses weight when immersed in the water, and

$$\frac{\text{Weight in air}}{\text{Loss of weight in water}} = \text{specific gravity.}$$

Unless great care be taken, the piece of coal used may be unduly contaminated with shale or pyrites, which would give too high a specific gravity, and several determinations on different pieces should be made, taking the mean or average result.

In cases where a large representative sample is taken, it is better to make the determination on the *powdered* sample, another method being, of course, necessary then. This is done in a *specific gravity bottle*, which is a small glass flask holding a known weight of water at a constant temperature (60° F.). A weighed portion of the powdered coal is placed in the flask, and water (at 60° F.) added until the flask is full, allowing a little time for the coal to become thoroughly wet and free from air bubbles. The stopper of the flask is then inserted, and any excess of water is forced through a fine hole specially made in the centre of the stopper, the flask carefully dried outside and weighed.

The coal has displaced its own volume of water, and therefore

$$\frac{\text{Weight of coal taken}}{\text{Weight taken + weight flask and water} - \text{flask, water, and sample.}} = \text{specific gravity.}$$

The specific gravity of coals ranges from about 1.2 to 1.4; and since 1 cubic foot of water weighs 62.5 lbs., the weight of a cubic foot of coal will range from  $62.5 \times 1.2$  to  $62.5 \times 1.4$  in lbs.

A bunker for storing coal measures 12 feet long, 10 feet wide, 8 feet high. The specific gravity of the coal to be stored in it is 1.25. How much will the bunker hold?

1 cubic foot water weighs 62.5 lbs.

∴ 1 cubic foot of this coal weighs  $62.5 \times 1.25 = 78.1$  lbs.

Cubic capacity of bunker =  $12 \times 10 \times 8 = 960$  cubic feet.

But if 1 cubic foot coal weighs 78.1 lbs.,

960 cubic feet coal weigh  $78.1 \times 960 = 74,976$  lbs.

∴  $\frac{74,976}{2,240}$  (lbs. in ton) = 33.47 tons (practically  $33\frac{1}{2}$  tons).

**The Moisture in Coal** is estimated by taking a weighed quantity in a copper dish, and exposing it to a temperature not exceeding 100° C. or 212° F. in a water bath, which is a small oven surrounded by a water jacket. The water is kept boiling by suitable means, and the temperature is constant and never exceeds the boiling point. The coal should be thus dried for eight or ten hours, depending on the amount taken, and also as to whether the sample is unusually wet or not.

It is advisable not to have the coal finely powdered for this test; the coarse sample should be used, otherwise there is a possibility of driving off volatile hydrocarbons as well. A too prolonged heating is undesirable, as the pyrites tends to oxidise, both effects giving too high a result for moisture. The loss in weight can be easily calculated into percentage. The safest method is to

expose a weighed portion of the coal under a glass bell jar in which sulphuric acid is placed, in a tray, for twenty - four hours. The loss in weight represents moisture.

**The Determination of Ash** is made by taking a weighed quantity, usually 1 or 2 grammes, and placing on a flat porcelain lid. (Platinum should not be used, as the carbon acts upon it in time, and causes it to crack and become brittle.) Heat *gently* at first over Bunsen flame or at a safe distance in front of the muffle furnace, and gradually increase the temperature until the sample can be safely placed inside the muffle at the front. Allow a plentiful supply of air, and finally place in a hotter part of the furnace for a few minutes. When all the carbonaceous matter has burnt away, take the dish out of the furnace, allow to cool, and weigh the ash. The percentage may be readily obtained thus—

2 grammes of coal left a residue of ash weighing .104 grammes: therefore 100 grammes would leave

$$\frac{.104 \times 100}{2} = 5.2 \text{ per cent.}$$

The colour of the ash varies from white to reddish brown, and often affords a valuable indication as to the contents of sulphur in the coal. A reddish-brown ash owes its colour to iron oxide, which denotes that iron *sulphide* or pyrites was present in the coal before burning; the sulphur has combined with oxygen of the air and gone off as a gas, but leaves the tell-tale iron *oxide* behind. A white ash, as a rule, denotes comparative freedom from sulphur in the coal from which it has been produced.

**Chemical Composition of the Ash.**—This varies with the kind of coal, and also with the district in which it is found; but all ashes contain silica, alumina, iron oxide, lime, magnesia, sulphuric acid (in the form of *sulphate*),

phosphoric acid (in the form of phosphate), potash and soda. The amounts usually vary as follows:—

Silica . . . . .	30 to 60	per cent.
Alumina and Oxide of Iron	35	„ 55 „
Lime . . . . .	5	„ 10 „
Magnesia . . . . .	traces	„ 2 „
Sulphuric Acid . . . . .	3	„ 8 „
Phosphoric Acid . . . . .	·5	„ 1·5 „
Potash and Soda . . . . .	2	„ 4 „

**The Volatile Hydrocarbons** are those compounds which are driven off at a full red heat when the fuel is heated *out of contact with air*.

A weighed quantity of fuel being taken (1 gramme coal, 2 to 5 grammes coke), in a weighed porcelain or platinum crucible, the cover of the latter is put on and the whole placed inside a larger crucible, and covered with coarsely powdered coke and the cover of the larger crucible, then put on over all (see Fig. 41). The fuel (if coal) is heated over a straight Bunsen flame, as long as any smoke is seen issuing from the crucible. It is then transferred quickly to a gas muffle at a red heat and the door closed to exclude air, and the heating continued from two to three minutes. Remove from muffle, cool under a bell jar (desiccator), and weigh loss of weight of inner crucible and contents. This loss represents the volatile matter, and the *residue* in the crucible is the *coke* produced. The appearance of the latter should be carefully noticed, as to whether it has coked well, or is hard or soft, etc.

The calculation is simple, *e.g.*:

1 gramme of coal, loses, on being thus treated ·387 grammes = 38·7 per cent. volatile and  $1·00 - ·378 = ·613 = 61·3$  per cent. coke produced.

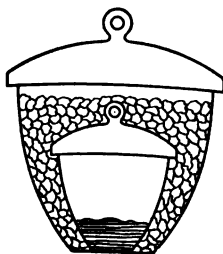


FIG. 41.—Estimation of volatile matter.

The volatile matter in a coke is estimated in the same way, but is, of course, much lower, varying from nil to 2·0 per cent.

**The Estimation of Sulphur** in coal or coke, for practical purposes, usually includes that from all three of the sources mentioned before—that is to say, the *total* sulphur.

There are several good reliable methods available for its estimation, but they require considerably more skill and care and knowledge of chemical action, than do those used for ash, moisture, or volatile determinations. Two methods, of different types, will now be described.

The principle in both depends upon converting the sulphur into a definite chemical compound (barium sulphate), the exact composition of which is known. It contains 13·73 per cent. of sulphur.

*Method 1.*—A weighed quantity of the coal or coke is intimately mixed with sodium carbonate and nitre (1 gramme coal + 3·5 grammes soda + 9·0 grammes nitre) and heated gently in a platinum crucible until fusion and chemical action begin. (Common salt is sometimes used instead of carbonate of soda.) The heating is continued until the mass becomes quiescent, and when cool the result is dissolved out in distilled water and transferred to a glass vessel (called a beaker.) Hydrochloric acid is then added until no further effervescence takes place and the solution is then boiled and filtered. All the coal or coke ought to have been oxidised. If any quantity remains on the paper, this should be washed and weighed, and its amount deducted from the original weight of fuel; or, better still, it should be fused with a little more fusion mixture and the result added to the first main portion.

The clear filtered solution is then boiled and barium chloride solution added to it. This reacts with the sulphate in solution (into which form all the sulphur has been converted), and a white substance (called a *precipitate*) appears

in the liquid, consisting of the above-mentioned *barium sulphate*. Thus—

Barium chloride + Sodium sulphate =

Barium sulphate + Sodium chloride.

A case of *mutual interchange*.

It remains then to allow this precipitate to settle out, to filter it off through filtering paper (Swedish quality is necessary), wash away all soluble matter by means of a *wash-bottle*, first using dilute acid and finally hot water alone, leaving the pure barium sulphate on the filter. The latter is burnt off in a platinum dish, and weighed (allowing from 1 to 2 *milligrammes* for the ash of the paper).

Then *e.g.* — If the barium sulphate produced from 1 gramme of the fuel weighs  $\cdot 123$  grammes, and the barium sulphate is known, beyond doubt, to contain 13·73 per cent. *sulphur*.

Therefore 13·73 per cent. of  $\cdot 123$  is the amount of sulphur in 1 gramme of fuel.

$\therefore 13\cdot73 \times \cdot 123 = 1\cdot68$  is the percentage.

*Method 2.*—This consists in heating the fuel in contact with magnesia and carbonate of soda or with lime, in an open dish, whereby the whole of the sulphur becomes sulphate of magnesia, soda or lime, as the case may be. The mass is transferred, after heating long enough to thoroughly decompose the fuel, to a beaker, dissolved in water and a slight excess of hydrochloric acid, with a few drops of bromine, filtered if necessary, any coal not acted upon being weighed, or reheated with lime or magnesia, the filtrate boiled and barium chloride added to it as before. The final operations being exactly the same as in method 1.

Method 2 is in very general use, and is to be preferred to the other for general purposes. In both methods the materials used must be themselves free from sulphur compounds, or these will cause too high a result. A blank test should be made, that is, by using all the materials *except the coal or coke*, and precipitating the

solution thus obtained with barium chloride and deducting the weight of any precipitate thus found from the other result.

**The Calorific Value or Heating Power of Coal** is determined by means of an apparatus known as a *calorimeter*, of which there are several modifications.

The principle of them all depends on either (1) burning the coal by oxidising chemical reagents, such as nitrates and chlorates, substances rich in oxygen; or (2) by burning the fuel in an *atmosphere* of oxygen gas. In both cases the heat produced is measured by delicate thermometers.

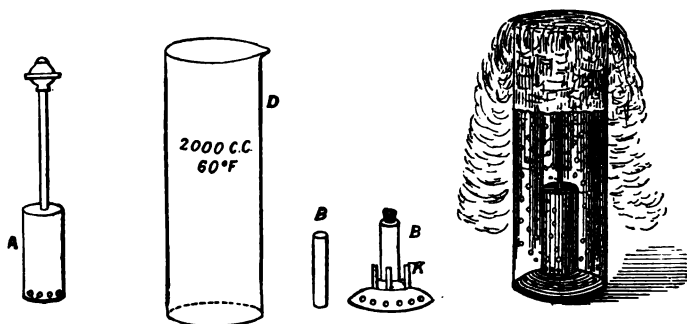


FIG. 42.—Calorimeter (showing details and calorimeter in action).

In the first-named method a weighed portion of the fuel is taken and intimately mixed with potassium nitrate and chlorate. The mixture is placed in a small copper crucible or tube closed at one end, and a piece of fuse inserted (lamp wick soaked in nitre solution, and dried).

The whole is then fixed on a special piece of apparatus which can be immersed in a measured quantity of water, without the water coming in actual contact with the contents of the crucible. The exact temperature of the water being taken by means of an accurate thermometer, the fuse is lit and the apparatus quickly immersed in the water. When the fuse gets down to the mixture of coal and oxidising agents, vigorous chemical action takes place,

due to the burning up of the coal by the oxygen of the nitrate and chlorate. The resulting gases, before they can escape, have to pass through the water surrounding the apparatus, and in doing so, give up their heat to it. When the combustion is complete, water is admitted to the inside of the apparatus, which is naturally very hot, and takes up the heat therefrom, so that the total heat evolved by the combustion of the coal is given up to the water. The apparatus itself absorbs some of the heat, and a *constant* allowance is made for this, usually 10 per cent. Knowing the volume of water and the number of degrees increase in temperature, caused by the burning fuel, it is possible to calculate its calorific value, and from this its *evaporative power*.

When oxygen gas is used, the apparatus is similar in construction, except that an appliance is fitted whereby the coal can be ignited by an electric current passing through a fine platinum wire, which becomes red-hot. Oxygen gas is led into the space surrounding the coal, instead of using the solid oxidising agents as described in method 1, and ignition once started by means of the hot wire, the fuel burns completely away, the escaping gases, as before, passing through the surrounding water, and giving up their heat to it.

In an actual test the temperature of water in the cylinder was  $14\cdot2^{\circ}$  C. before the combustion and  $21\cdot3^{\circ}$  C. after.

The increase in temperature therefore was  $7\cdot1^{\circ}$  C.

2 grammes of coal were used, and the volume of the water was 2000 cubic centimetres.

On adding 10 per cent. for absorption of heat by the apparatus, the actual increase in temperature was  $7\cdot1^{\circ} + \cdot71^{\circ} = 7\cdot81^{\circ}$  C.

This multiplied by 1000 = heat units or calories = 7810,

$$\text{and } \frac{7810}{537} = 14\cdot54 \text{ lbs. of water}$$

converted into steam by 1 lb. of the fuel (537 being the latent heat of vaporisation of water).

In the more elaborate apparatus known as the *bomb calorimeter*, the oxygen gas is supplied to the fuel in a compressed form, and the combustion takes place under several atmospheres' pressure. The combustion is started electrically, and the bomb has to be very strongly made to withstand the high pressure of the oxygen, increased by the temperature of the combustion. The bulk of metal employed necessitates a careful correction for the heat absorbed by the apparatus itself.

The calorific or heating value of fuel can be calculated if its chemical composition be known, but the theoretical results thus obtained are always higher than the results obtained in practice. Radiation and absorption of heat partially account for this, as do also certain chemical actions which require heat for their completion.

When carbon burns completely to carbon-di-oxide it produces 8080 heat units, and every 1 lb. of carbon requires 11·57 lbs. of air.

Hydrogen, when burnt to form water, produces 34500 heat units, and 34·7 lbs. of air are required for its combustion.

Sulphur, when burnt, produces 2220 heat units, and every lb. requires 4·35 lbs. of air.

Carbonic oxide, burning to form carbon-di-oxide, generates 2403 heat units, and requires 2·48 lbs. of air per lb. of gas.

A calorie, or heat unit, is the quantity of heat required to raise 1 kilogramme of water 1° C., and a British Thermal unit is the amount of heat required to raise 1 lb. water 1° F., hence 1 calorie is equal to 3·968 British Thermal units.

The increase in temperature of the water is usually taken as being from 0° C. to 1° C., but so long as the temperature is below 100° *finally*, it is not material. After what has been said about latent heat (which see), it would be absurd to take the amount of heat required to raise the temperature of water from 100° C. or 212° F.

(its boiling point) to  $101^{\circ}\text{C.}$  or  $213^{\circ}\text{F.}$ , because we know that unless the pressure be increased or solids added to the water, that the latter is not increased in temperature once the boiling point is reached.

If the increase in temperature be measured in Fahrenheit degrees, then the numbers representing the calorific powers, as expressed by Centigrade degrees, must be increased in the proportion of 9 : 5.

The heating power of a fuel depends upon the carbon and hydrogen *which are available* for combustion.

If any oxygen be present (and there always is some proportion in coals), that oxygen is taken to be already combined with its share of hydrogen, and therefore is useless for heat production when the fuel is burnt. It has been shown already that oxygen is combined with  $\frac{1}{8}$ th its weight of hydrogen to form water, and when a fuel containing oxygen burns, it uses up this proportion of hydrogen, and this is always allowed for in calculating the heating power of a fuel from its analysis.

For instance, assume an analysis of coal to show—

Carbon . . . .	80 per cent.
Hydrogen . . . .	5 „
Oxygen . . . .	8 „
Nitrogen . . . .	2 „
Ash . . . .	5 „

Then 100 lbs. of this fuel contain 80 lbs. of carbon, 5 of hydrogen, and 8 of oxygen.

The 8 lbs. of oxygen will require 1 lb. of hydrogen, and this leaves 4 lbs. hydrogen available for heat production, and the calculated heating power of this fuel will be—

Heat due to carbon . . .	$8080 \times 80 = 646400$
„ hydrogen . . .	$34500 \times 4 = 138000$

Total heat units . . .	784400
------------------------	--------

$\therefore$  1 lb. of this fuel would generate  $\frac{784400}{100} = 7844.00$

heat units.

A certain deduction would have to come off this again, for the heat which becomes latent in converting the water (formed by the union of the oxygen and hydrogen in the fuel) into steam.

If the coal be wet, then a further amount of heat is lost in converting this moisture into steam—

			Centigrade heat units produced.	Lbs. of air required by 1 lb. of the substance.
Carbon, when burnt to	CO <sub>2</sub>		8080	11·57
Hydrogen, „	H <sub>2</sub> O		34500	34·7
CO, „	CO <sub>2</sub>		2403	2·48
CH <sub>4</sub> , „	CO <sub>2</sub> and H <sub>2</sub> O		13063	17·35
Sulphur, „	SO <sub>2</sub>		2220	4·35

## CHAPTER XII.

### STRATA ADJOINING THE COAL MEASURES.

THE strata adjoining the coal measures consist of shales and clays of various kinds, clay ironstones, ganister, etc., more or less valuable for working purposes according to their composition.

Chemically considered, clay is a *compound* of silica (the chief constituent of sand), with alumina (the oxide of the now well-known metal, aluminium), and *combined* water. In other words, it is a hydrated silicate of alumina. In addition to these substances, there are always varying amounts of iron oxide, lime, magnesia, and the alkalies, potash and soda.

All clays have been formed by the decomposition of felspathic rocks. These are silicates of alumina, potash and soda. When exposed to the weathering action of rain, wind, frost, running water, carbon-di-oxide and oxygen, they gradually decompose, and certain of their constituents (sodium and potassium) become carbonates, which are soluble in water, and therefore are washed away. This leaves the silicate of alumina in a more or less pure condition, often contaminated with oxide of iron, but the purer clays (*e.g.* china clay) are free almost from this constituent. The colour of a clay is a good indication of the amount of iron it contains, a reddish-brown clay being richer in iron than one of a paler colour. China clay is almost white.

**Suitability of Clays for various purposes.**—The clays are suitable for various purposes according to their composition; that which would be satisfactory for making

building bricks would be absolutely useless for making a refractory brick, that is one which will withstand intense heat, such as that required in metallurgical processes.

A good clay for refractory purposes should contain 50 to 60 per cent. silica and from 20 to 30 per cent. alumina; there should be very small amounts, if any, of lime, magnesia, and especially alkalies, potash and soda. Small amounts of iron may be present, but should not exceed 2 or 3 per cent. of oxide. Iron in the form of pyrites is very objectionable. Lime, magnesia, potash, and soda all tend to lower the fusing or melting point of the brick, since they form readily fusible silicates with the silica. Potash and soda are especially objectionable on this account. In clays to be used for building-bricks, the presence of these impurities is not so serious. The oxide of iron present determines the colour of the brick, and much more lime, magnesia, and alkalies may be tolerated than in a fire-clay.

It has been previously stated that clays contain *combined* water, usually about 10 per cent. To this they owe their *plasticity*, that is the property of being moulded into various forms. If this combined water be driven off by heat, the clay never regains its plastic power. A practical method of testing the refractory nature of a clay is to mould a portion into the shape of a rectangle, or cube, or pyramid with *sharp edges*; after carefully and slowly drying, the test piece is subjected to a high temperature. A clay of high refractory quality ought not to show the slightest rounding-off or fusing of the sharp edges.

Many things have to be considered in dealing with refractory bricks, especially the purposes and uses to which they are intended to be put. For instance, a brick may consist principally of silica (silica brick), containing as much as 98 per cent. pure silica, and such a brick subjected *alone* to a very intense heat would be unaltered, practically. If, however, a slag containing much lime, or if lime itself or ferrous oxide be brought into contact with the brick, fusible

compounds, silicates of lime or iron, would be formed, and cause the brick to "run." Similarly, lime and magnesia *alone* are infusible, even at the most intense temperatures, but if they come into contact with siliceous substances they melt fairly easily.

The following are analyses of refractory materials:—

	Fire-clay.	Fire-clay.	Fire-brick.	Magnesite Brick.	Silica Brick.
	Per cent.	Per cent.	Per cent.	Per cent.	Per cent.
Silica . . .	63·80	50·20	58·00	2·40	96·80
Alumina . . .	22·20	31·71	34·8	·50	·60
Iron Oxide . . .	2·97	2·78	3·8	·40	1·50
Manganese Oxide	traces.	traces.	traces.	—	—
Lime . . .	1·00	1·00	1·0	·10	·80
Magnesia . . .	traces.	traces.	1·4	95·8	·20
Alkalies, Potash and Soda . }	·52	·91	1·0	not estd.	·50
Combined water and organic matter . }	9·50	13·40	nil.	·50	nil.

**Aluminous Shales.**—The shales are also compounds of silica and alumina, and they are often treated for the alumina they contain.

Alum shale is a coaly mineral, containing pyrites (bisulphide of iron). On roasting this shale and then "weathering," the sulphur of the pyrites becomes partially converted into sulphuric acid, which acts upon the silicate of alumina, forming sulphate of alumina. The mass is washed with water (lixivated or leached) to get out all the soluble matter, and potassium sulphate is added in sufficient quantity to form ordinary *alum*, which is a double sulphate of aluminium and potassium. Certain iron compounds are formed at the same time, but being more soluble in water they remain dissolved, while the less soluble alum is allowed to crystallise out in vats.

Some bituminous shales are subjected to distillation to

recover oils from them, and most of them are now worked for ammonia as well, and this latter forms an important branch of the shale industry. The principle underlying this recovery is practically the same as that referred to under by-product recovery in coking.

The shales richest in hydrocarbons probably contain more vegetable matter than those which are rich in ammonia; the latter probably contain a greater proportion of animal remains.

**Iron Pyrites** has been referred to several times previously. It is commonly met with in and near the coal measures. It is a compound of iron and sulphur, and is often used as a source of sulphur in various manufacturing processes, *e.g.* in sulphuric acid making, the residue (which is chiefly oxide of iron) being then used for the iron which it contains. The pyrites is also, in some cases, converted into "copperas" or "green vitriol" by weathering. The name "copperas" for the sulphate of *iron* produced is somewhat misleading. The sulphur of the pyrites takes up oxygen forming sulphate and at the same time *free* sulphuric acid is produced (see acid waters); this free acid is neutralised and converted into more sulphate of iron by the addition of scrap iron or steel. The salt produced is crystallised out in vats.

**Ganister** is a rock met with in certain districts near the coal measures. It is almost entirely silica (at least 95 or 96 per cent.), and is very refractory, being consequently used for purposes where great heat has to be withstood, such as lining the Bessemer converter in which steel is made by the acid process. An analysis of a brick made of ganister was as follows:—

Silica . . . .	96·0
Alumina . . . .	1·0
Iron Oxide . . . .	·9
Lime . . . .	1·2
Magnesia . . . .	·6

**Ironstones.**—The strata near the coal measures often contain workable amounts of iron, generally in the form of carbonate. They are sometimes associated with silica and alumina, and usually a fair amount of carbonate of lime, with phosphoric acid and sulphur.

The analyses given below show the differences between the various kinds of ironstones, and especially the composition of those associated with the coal measures. The iron in the latter, it will be noticed, is mainly in the form of carbonate, while others have their iron principally in the form of peroxide.

It is generally advantageous to calcine (that is, heat strongly in a kiln) ironstones in which the iron is as carbonate, before smelting. This drives off carbon-di-oxide and combined water, and the final result of such heating, in the case of the "cockle-shell" bed (analysis No. 8), would be to more than double the amount of metallic iron (29.47 per cent.), the actual loss due to heating being 51.3 per cent. Clay ironstones vary in colour from pale brown to black, the latter caused by coaly or carbonaceous matter mixed with them, sometimes in sufficient quantity to enable them to be smelted without the use of further fuel.

ANALYSES OF IRONSTONES (TEN VARIETIES).<sup>1</sup>

	1	2	3	4	5	6	7	8	9	10
	Per cent.	Per cent.	Per cent.	Per cent.	Per cent.	Per cent.	Per cent.	Per cent.	Per cent.	Per cent.
Silica . . . . .	1.00	8.40	3.60	3.00	7.32	17.80	22.1	1.90	10.75	28.30
Alumina . . . . .	.50	6.00	1.0	4.0	5.85	9.50	2.25	1.00	4.62	10.10
Iron peroxide . . .	98.1	...	...	3.2	43.27	44.50	.64	2.86	...	...
Iron protoxide . .	nil.	44.50	54.0	45.0	...	...	38.50	15.80	18.97	30.47
Manganese oxide . .	.25	2.10	1.5	1.7	1.40	.75	...	1.30	...	.93
Lime carbonate . .	trace	2.05	traces	5.0	31.60	15.00	1.4	29.60	35.71	1.80
Magnesia carbonate .	trace	3.51	traces	.75	traces	trace	1.93	trace	5.81	9.20
Phosphoric acid . .	trace	.20	.10	.05	1.15	.86	.30	.42	.42	...
Sulphuric acid . .	nil.	2.10	3.2	.65	.10	...	.60	.65	.50	...
Organic matter and combined water . . . . .	nil.	4.0	4.0	8.0	...	nil.	trace	30.37	14.5	20.60
Iron carbonate . .	nil.	71.80	87.0	72.5	...	...	62.0	25.47	30.5	49.10
Metallic iron . . .	68.70	3.40	42.0	35.0	30.3	31.13	30.0	14.35	14.76	23.70

<sup>1</sup> See next page.

The ironstones referred to in the foregoing Table are the following:—

1. Purest hematite ore from Cumberland district.
2. Clay ironstone.
3. Black clay ironstone from coal measures.
4. Blackband ironstone.
5. Lincolnshire ore.
6. Fossiliferous stone from Lincoln ore.
7. Material found between two beds of coal near St Helens, Lancashire.
8. Cockle-shell bed, above the Arley Mine.
9. Dark shale from cockle-bed, 67 yards above the Arley Mine. Full of shells.
10. Material found under fire-clay, below yard coal.

**Granite.**—This rock is so called because it is readily seen to be made up of small *grains*, and it consists of three minerals: quartz, felspar, and mica. Quartz is almost pure silica. Felspar and mica are compounds of silica with alumina iron oxide, lime, magnesia, potash and soda. By “weathering,” these constituents become disintegrated, and gradually crumble away; the quartz grains thus freed go to form sand, and the silicates of alumina, etc., from the felspar and mica, become converted into the various clays, as already described.

## CHAPTER XIII.

### MAGNETISM AND ELECTRICITY.

It is altogether beyond the scope of this book to presume to deal in detail with the various applications of electricity in connection with mining operations. The purpose, in adding a few notes on elementary magnetism and electricity, is to explain in as short a manner as possible the principles underlying the more important operations, and thus enable the student to more readily understand the appliances.

**Magnetism and its Effects.**—The natural magnet commonly known as a lodestone is one of the oxides of iron, and is, in fact, one of the purest forms in which iron ore is found. It has the power of attracting or drawing iron or steel towards itself. Further, if a small lump of lodestone be suspended by a fine string, it will turn so as to place itself in a position pointing magnetic north and south. There are objections to the use of the natural magnet in practice, and artificial magnets are made of steel bars and rods. When hard steel is rubbed with a piece of lodestone or with another magnet, or has a current of electricity passed round it by means of a coil of wire, the steel takes up magnetic properties *permanently*—that is, becomes magnetised.

A small strip of steel, suitably balanced on a point and magnetised, becomes a magnetic needle, and is used in a variety of ways. The end of the needle pointing north is usually marked in some way. On bringing a bar magnet to one end of this needle, the latter is either attracted or repelled, according to the "poles" which

are near each other : *unlike poles attract, like poles repel each other*. The action of magnetism goes on through substances which are not magnetic. By holding a piece of card or stiff paper above a bar magnet, and sprinkling iron filings on the paper, the filings arrange themselves in well-marked lines, which are taken to indicate the lines of magnetic force. Although one end of a bar magnet has north-seeking magnetism, and the other south, it is impossible to break a bar so as to isolate one from the other. However small the pieces be broken, each one is still a small magnet, or is said to have *polarity*.

Though the magnetic needle points north and south, it is found that the direction does not quite coincide with the geographical north and south. The difference varies from time to time, and this difference is spoken of as the "declination" of the needle. Again, if a needle be balanced so as to swing *vertically*, and *then* magnetised, it is found to be no longer truly balanced, but to incline or dip downwards with its north-seeking end. This is spoken of as the "inclination" or "dip" of the magnetic needle, and (like the declination) varies from time to time. Fig. 43 illustrates a dipping needle.

#### **Chemical Properties of the Electric Current.—**

When a plate of pure zinc or a plate of amalgamated zinc is dipped in dilute sulphuric acid there is no action, and similarly with a plate of copper or platinum foil; but if the portions of the metals outside the liquid be joined by a wire, bubbles of hydrogen gas begin to come off from the platinum or copper plate, and if the wires from the plates be attached to a galvanometer or current indicator, it indicates the passage of a current. The chemical energy has been transformed into electrical energy. Similar reactions take place between couples of other metals when a current of electricity, sufficiently strong, is passed through acidu-

lated water. The latter is decomposed or split up into the two gases which compose it—oxygen and hydrogen.

This effect of the current is called *electrolysis*, and takes place not only with water, but with many other solutions. The metals potassium and sodium were discovered by electrolysis, and this principle underlies many manufacturing operations at the present day—notably the production of the wonderful metal aluminium.

The operation of electroplating is one of electrolysis, the gold or silver or nickel solution, etc. being decomposed by the current and its metal deposited on the object to be coated, the metal being supplied, as fast as deposited, by a strip of the same kind of metal, suspended in the solution, and connected to one end of the battery, the object to be plated being attached to the other end. Electrotyping is conducted on the same principles.

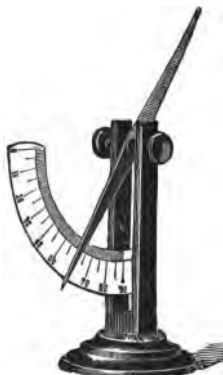


FIG. 43.—Dipping needle.

**Magnetic Effects of a Current.**—A current of electricity passing through a wire causes that wire to have magnetic properties, and if such a wire be held over a magnetic needle, the latter is deflected or turned aside, and takes up a position almost at right angles to its original one. Instruments applying this principle are used to recognise the presence of electric currents. They are called *galvanometers*, and consist of two needles, one above the other, the lower one surrounded by coils of insulated copper wire (copper wire is *insulated* by covering it with silk, cotton, or gutta-percha, so as to prevent the current leaking from one turn of wire to another). The coils of wire multiply the action of the current by *induction* (see p. 133), and by means of a

pointer attached to the needle-support, indicate when a current is passing through the galvanometer.

When a length of wire conveying a current is coiled round a piece of iron or mild steel, the latter becomes magnetised. On cutting off the current the magnetism at once ceases. If a hard steel has been used the magnetism remains, or becomes permanent.

**Effects of Currents upon each other.**—Currents passing through wires have certain effects on each other.



FIG. 44.—Solenoid.

If they run alongside of one another in the same direction, they attract one another; but if in opposite directions, repel one another.

A coil of wire wound in form of a helix, through which a current passes, acts just like a bar magnet, and has north and south polarity, also setting itself, if free to do so, in the direction of magnetic north and south. Such a coil of wire is called a *solenoid* (Fig. 44).

An ordinary bar magnet held near the end of a solenoid attracts or repels it according to the poles, just as another magnet would be attracted or repelled, and one solenoid will attract or repel another solenoid.

**Electro-Magnets** are usually of the horse-shoe shape, in one piece, or consist of two rods joined at the ends by a strap of iron. They should be made of *soft iron* rods, or better, of bundles of soft iron wire. Very powerful electro-magnets can be made, their strength depending, up to certain limits, upon the number of coils of wire round them, and also upon the strength of the current passing through the wire (see Figs. 45 and 46).

**Use of Electro-Magnets.**—The principle of the electro-magnet is applied in many ways—for instance, in

electric bell construction, in the magnetic concentration of iron ores, in the construction of induction machines, etc.

In the electric bell (Fig. 47), the striker is so arranged

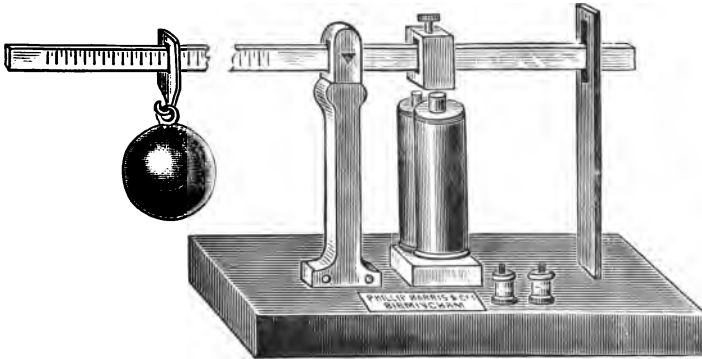


FIG. 45.—Electro-magnet.

that a small electro-magnet attracts the rod carrying the striker towards the poles of the magnet, and in doing so cuts off the electric current automatically. The current being cut off, the electro-magnet is no longer magnetic and the striker falls back to its original position, when contact is again made, and the process repeated. This goes on very rapidly, and causes the bell to be rung by the striker. Locks for safety lamps have been devised, which require the use of a strong magnet to unlock them. Such lamps cannot, of course, be opened without a special appliance, and are therefore very safe in that respect.

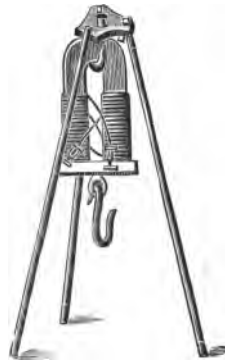


FIG. 46.—Electro-magnet.

**Primary Batteries.** — The batteries supplying the electrical energy for purposes of bell-ringing, signalling, telephony, and so on, depend on

chemical principles, and are generally primary batteries. We have seen that a battery or cell of a battery under simple conditions has one of its plates coated with bubbles of hydrogen gas. This gaseous layer causes a *back current*, and therefore tends to destroy the effect of the primary current. The main object of most of the

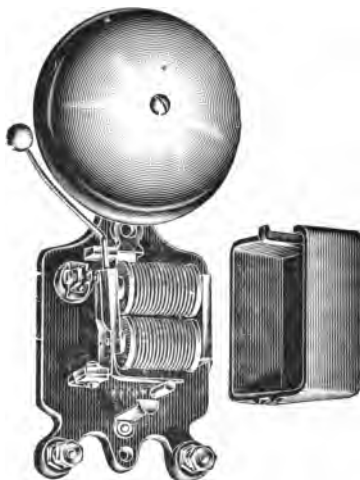


FIG. 47.—Electric bell.

chemical agents used in primary cells is to neutralise or destroy this hydrogen gas.

**Primary Cells in common use.**—The cells in most general use are either of the Leclanché (Fig. 48), Daniell, or bichromate type (Fig. 49.) In the Leclanché cell, the liquid used is a solution of sal-ammoniac, or ammonium chloride, and the agent used to destroy the hydrogen, or the *depolarising* agent, is peroxide of manganese. In the Daniell cell, copper sulphate is used for the same purpose, metallic copper being produced. In the bichromate cell, the depolarising agent is chromic acid, which gives up its oxygen and destroys the hydrogen.

With other cells, not often used for bell or signalling work, the depolarising agent is nitric acid (Bunsen's and Grove's cells) (Fig. 50).

**Induction.**—It has been shown that a coil of wire through which a current is passing, passed round an iron or steel bar, causes the latter to become magnetised. Now the converse also applies, namely, that if a permanent magnet be pushed into a coil of wire (Fig. 51) a *momentary* current of electricity is produced in that wire, and further, a *momentary* current is also produced when the bar is withdrawn. These are spoken of as *induced* currents, and have many important practical applications. Though these currents are momentary or practically instantaneous, they have all the properties of ordinary currents.



FIG. 48.—Leclanché cell.

In the **Induction Coil** (Fig. 52), a small original primary current is built up by induction into a powerful secondary current. As shown in the illustration, the coil consists of two necessary parts—a primary coil (the thin cylinder) and a secondary coil (the outer cylinder); the illustration shows the inner coil pulled out, but when in use its position is inside the larger coil. The primary wire is only a few feet in length, while the secondary may be at least a mile in length, each layer or strand of wire being insulated from its neighbour by paraffin wax or shellac varnish. By means of a “make and break” contact, like that on an electric bell, a rapid series of currents is sent along the primary wire; these induce currents in the secondary wires, and the number of coils multiplies the effect. The induced current is collected from the terminals of the secondary wires, and may be used for exploding gases or detonators, or displaying vacuum tubes, etc., and bringing about chemical decomposition.

**The Principle of the Telephone.**—The telephone depends for its properties upon induction. A permanent magnet has on one end of it a bobbin containing a coil of fine wire, the ends of which are connected in a closed electrical circuit. Opposite this end of the magnet is fixed a thin plate of soft iron (ferrotype). This disc or diaphragm

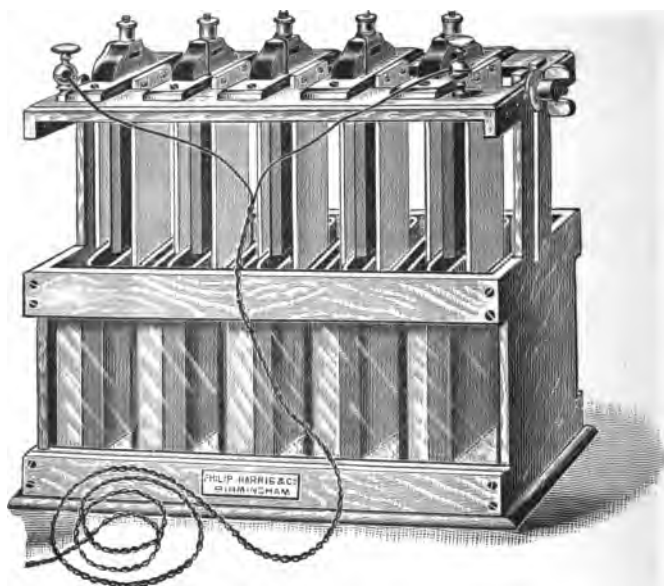


FIG. 49.—Bichromate or Daniell battery.

is magnetised by the current passing through the coil on the bobbin. The disc being fixed in the mouth of the telephone vibrates when spoken into, so that it moves to and fro, the movements varying with the sound which produces them. The to and fro movement causes induced currents to be set up and produce alterations in the magnetism of the permanent steel magnet, and hence currents are set up also in the coil of wire, which are transferred to the receiving end of the telephone, where similar

effects are produced in the reversed order, the currents in this case causing the disc to vibrate and reproduce faithfully every sound.

**Magneto - Electric Machines.**—When a conductor moves in a *magnetic field*—that is, in an area in which lines of magnetic force are acting—currents are produced in the conductor. Upon this principle, most magneto - electric



FIG. 50.—Bunsen or Grove's cell.



FIG. 51.—Illustrating induction.

machines are constructed. These machines are used in shot firing, and are usually constructed so that a number of coils of insulated copper wire, suitably wound, are made to revolve before the poles of a powerful permanent magnet.

**Dynamo Construction.**—Permanent steel magnets, as used in magneto-electric machines, are supplanted by soft iron or electro-magnets. A small amount of residual magnetism, always present in the soft iron, is found to

generate a small amount of electricity in the revolving coils of wire (called the *armature*); and this induced current, acting upon the electro-magnet, increases its magnetism, and therefore increases the number of lines of force in the magnetic field. The current is thus very rapidly multiplied to a maximum, depending up to certain limits upon the speed at which the armature is revolving, upon the number of coils on the armature, and upon the intensity of the magnetic field. Such a machine is

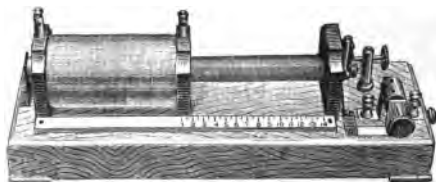


FIG. 52.—Induction coil.

called a *dynamo*, and is therefore a machine for changing mechanical work into electrical energy.

In an electric *motor*, the electric energy derived from a dynamo is changed again into mechanical energy. A dynamo and motor are therefore the opposite of one another. In the former the armature is made to revolve by mechanical energy in a magnetic field; in the latter the magnetic field, produced by current from the dynamo, causes the armature to revolve, and be available for mechanical work.

**Secondary Batteries or Accumulators.**—We have seen that when a current of electricity is sent through water, the latter is decomposed into its elements hydrogen and oxygen. If these gases be collected separately in tubes containing a platinum plate or *electrode*, under certain conditions a reverse current is established, and a battery has been devised by Groves on these principles called a gas battery.

Upon these lines most secondary batteries or accumulators are constructed (Fig. 53). They all involve chemical action, and one of the commonest forms consists of leaden plates in dilute sulphuric acid. On passing a current into a cell containing two such plates, electrolysis is set up; oxygen is liberated at one plate and converts the metallic

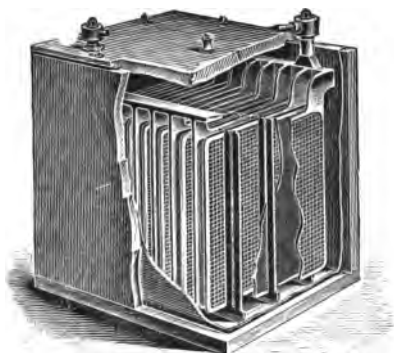


FIG. 53.—Secondary battery accumulator.

lead into peroxide of lead; at the other plate hydrogen is liberated. After a time the current is reversed, then the lead peroxide becomes *reduced* by the hydrogen which is now liberated at that plate, while the other plate becomes coated with peroxide. This is carried on for some time with occasional changing of the current, and finally the plates remain, one coated with *spongy* lead, the other with peroxide. The spongy condition of the metal is one of the necessary features of the cell, and it is brought about by the reduction of lead peroxide. No mechanical means could prepare it in such a suitable condition. It is now found to be better to give the leaden plates a preliminary coating of lead oxide before passing the current.

The secondary cell thus produced (or series of cells forming a battery) may be charged from a dynamo, or from a primary battery, as required. The Sussman electric lamp

is an example of one mounted on a secondary battery. It will maintain its lighting power for eight or nine hours, and is practically the same weight as an ordinary oil safety lamp. It is recharged as required from a dynamo.

**Lighting Effect of the Electric Current.**—As regards the lighting effect of the current, two forms of lamps are, broadly speaking, in general use, the incandescent and the arc. Both primary and secondary batteries are used as a source of the chemical energy, which is transformed into light in the case of the incandescent lamp. As a rule primary batteries are unsatisfactory for this purpose, and most secondary batteries have the disadvantage of weight. No doubt this difficulty will be eventually overcome.

The *incandescent* light is caused by the passage of a current of electricity from a primary or secondary battery, or from a dynamo, through a fine filament or thread of carbon. At one time these filaments were made from bamboo fibre, which had many imperfections and disadvantages. Many improvements have been effected, and the filament is now usually made by dissolving cellulose in chloride of zinc, which produces a semi-solid substance. This is "squirted" or forced through small holes, and in this form is very like hard white of egg in appearance. By heating this material out of contact with air a filament of carbon is left, and this is suitably mounted in a glass globe, which is then emptied of air. The resistance of this carbon thread to the passage of the current causes it to become white-hot. It is, of course, absolutely necessary that no air has access to it, otherwise the filament would burn away almost instantly. Incandescent lights are usually considered one of the safest methods of lighting in the mine, but under some conditions may be dangerous.

Messrs Couriot and Meunier (*Comptes Rendus*, 1898) experimented with an explosive mixture of fire-damp and air (9.5 per cent. methane) and 10-candle power lamps run

with a voltage of 15 to 110 volts; no explosion could be produced provided the filament of the lamp remained unbroken. Water and carbon dioxide were produced, and only on one occasion was there explosion. On investigating this case it was found that the filament was broken, and explosion was probably caused by sparking between the ends of the broken filament. Other attempts to produce an explosion were unsuccessful.

Heise and Theim found, when experimenting with a mine gas in Westphalia (1898) with a 500-volt current, that the gaseous mixture always exploded when the glass was carefully removed and the filament exposed. When the glass was broken the gas was sometimes ignited, but not always. With arc lamps and with sparks from machines the conclusions arrived at were that all visible naked sparks are dangerous. In mines lighted by glow lamps precautions should be taken to ensure that the filament gets *completely* broken if the lamp-glass breaks.

The Nernst incandescent electric lamp is one in which the filament, instead of being enclosed in a vacuous flask or globe, glows in the open air. The filament consists of certain rare earths, similar to those used in the manufacture of incandescent mantles—namely, thoria and zirconia. These substances are what are called “secondary conductors,” that is, they require to be heated before an electric current can be passed through them. The heating is done by heaters in the form of hollow cylinders made of kaolin, or best fire-clay, round which platinum wire is wound. When the current is first turned on it passes through the heater, which imparts its heat to the filament, until the latter is hot enough to allow of the passage of the current through it, causing incandescence. The current is then cut off automatically from the heater.

Ordinary *arc* lights consist of two pointed carbon rods fitted in suitable machinery whereby they are automatically kept at a certain distance from each other. The regulator of this distance is usually worked by an electro-

magnet. When the points of the rods are placed together the current meets with great resistance at these points, and makes them white-hot; on separating them slightly a brilliant light continues to be given out.

**Electrical Welding.**—The welding together of metals by means of electricity is a very useful application of the heating effect that the current possesses. The operation can be made use of not only with the metal iron, but with others, such as brass, copper, zinc, etc., and can be applied not only to metals of the same kind, but with others—for example, brass to tin or copper, etc.

The process can be used when ordinary methods of welding would not be possible. Thus as in the case of cracked portions of machinery, such as fly-wheel arms, etc., the cracks may be filled with filings or fine drillings, and the welding done on the spot. This is a great advantage, especially when the machinery is underground, and when, if ordinary welding had to be done, the part would have to be removed to the surface.

It would not be advantageous to erect a generating plant especially for welding purposes, but where current is already being used, say coal cutting, the necessary addition for welding purposes would not be a very serious item. As most large modern collieries are using electricity extensively, no doubt electrical welding will be more generally used before long.

**Chemical Welding.**—In connection with the subject of welding, it is worthy of note that a chemical method is in use which can be applied *in situ* also.

The principle depends upon the fact that when oxide of iron is intimately mixed with powdered or granulated aluminium, under suitable conditions a very intense chemical action takes place, which develops an enormous amount of heat. The aluminium is oxidised at the expense of the

iron oxide, and the heat is the result of the chemical change.

Any portion of machinery or structural work which it is desired to weld has a suitable amount of the mixture (called *thermite*) placed round the part to be welded, and then fired. It is found that it is sufficient to start the chemical action at one spot only, by means of a kind of fuse, really a mixture of barium peroxide and aluminium, which is fired by a short piece of magnesium wire; once commenced, the action spreads throughout the whole mass of thermite, and raises the material which it surrounds to a welding heat. The process was devised by Goldschmidt.

## APPENDIX.

### I. USEFUL TABLES, ETC.

Conversion of thermometer degrees from one scale to another.

C.° to F.°—Multiply by 9, divide by 5, add 32.

F.° to C.°—Subtract 32, multiply by 5, divide by 9.

Rough estimation of temperatures—

Just glowing in the dark, about 525° C.

Dark red	.	.	„	700°	„
Cherry red	.	.	„	910°	„
Bright cherry red	.	.	„	1000°	„
Orange	.	.	„	1160°	„
White	.	.	„	1300°	„
Dazzling bluish white	.	„	1500°	„	
Electric arc	.	.	„	3500°	„

Area of circle	.	=	radius squared × 3·1416
Area of square	.	=	length of one side squared
Volume of a sphere	.	=	radius cubed × 3·1416 × $\frac{4}{3}$
Surface of sphere	.	=	diameter squared × 3·1416
Volume of a cylinder	.	=	area of base × height
Volume of a cube	.	=	length of edge, cubed
One atmospheric pressure	=	{	14·7 lbs. per square inch, or 2116 lbs. per square foot

Specific gravities of common metals—

Platinum	.	.	21·5	Mercury	.	.	13·6
Gold	.	.	19·3	Lead	.	.	11·4

Silver . . .	10·5	Chromium . . .	7·3
Copper . . .	8·9	Zinc . . .	6·9
Nickel . . .	8·8	Aluminium . . .	2·7
Manganese . . .	8·0	Magnesium . . .	1·75
Iron . . .	7·8	Potassium . . .	·86
Tin . . .	7·3		

## Melting points—

Aluminium . . .	700° C.	Magnesium . . .	500° C.
Copper . . .	1100° „	Mercury . . .	−39° „
Gold . . .	1075° „	Nickel . . .	1500° „
Cast iron (white)	1080° „	Platinum . . .	1775° „
Cast iron (grey)	1275° „	Silver . . .	960° „
Wrought iron . .	1550° „	Tin . . .	230° „
Steel . . .	1375° „	Zinc . . .	412° „
Lead . . .	326° „		

## Common names of chemical substances—

Chemical name.	Common name.
Potassium, aluminium sulphate.	Alum.
Copper sulphate.	Blue vitriol or blue stone.
Iron sulphate.	Copperas or green vitriol.
Zinc sulphate.	White vitriol.
Mercurous chloride.	Calomel.
Mercuric chloride.	Corrosive sublimate.
Nitric acid.	Aqua fortis.
Nitro-hydrochloric acid.	Aqua regia.
Phenol.	Carbolic acid.
Magnesium sulphate.	Epsom salts.
Sodium or potassium hydrate.	Caustic soda or potash.
Bisulphide of iron.	Coal brasses, iron pyrites.
Tin binocide.	Putty powder.
Iron per oxide.	Rouge.
Silver nitrate.	Lunar caustic.
Sulphuric acid.	Oil of vitriol.

Chemical name.	Common name.
Ammonium chloride.	Sal-ammoniac.
Potassium nitrate.	Nitre or saltpetre.
Calcium hydrate.	Slaked lime.
Calcium oxide.	Quicklime.
Calcium sulphate.	Plaster of Paris.
Calcium carbonate.	Chalk, marble, limestone.
Sodium sulphate.	Glauber's salt.
Acetate of copper.	Verdigris.
Ammonia.	Spirits of hartshorn.
Hydrochloric acid.	Spirits of salts or muriatic acid.
Nitrous oxide.	Laughing gas.
Sodium carbonate.	Soda.
Acetic acid (dilute).	Vinegar.
Sodium chloride.	Common salt (table salt).

## English equivalents of Metric System units—

1 Metre	= 39·37 English inches.
1 Kilogramme (kilo)	= 2·20 pounds.
1 Litre (1 cubic decimetre)	= 1·76 pints (61·02 cubic inches).
1 Gramme	= 15·43 grains.

1 Gallon of water weighs 10 lbs. = 70,000 grains.

∴ 224 Gallons weigh 1 ton.

## Symbols and specific gravities of common gases—

Name.	Symbol.	Spec. gravity, air being 1.
Hydrogen	H	·06926
Marsh gas	CH <sub>4</sub>	·5592
Carbon-di-oxide	CO <sub>2</sub>	1·529
Carbonic oxide	CO	·9678
Sulphuretted hydrogen	H <sub>2</sub> S	1·1748
Sulphur-di-oxide	SO <sub>2</sub>	2·247
Oxygen	O	1·1056
Nitrogen	N	·9713

Coefficients of linear expansion between 0° and 100° C.

Brass	.	.	.	.	·00001878
Gold	.	.	.	.	·00001466
Copper	.	.	.	.	·00001718
Iron	.	.	.	.	·00001220
Steel	.	.	.	.	·00001079
Lead	.	.	.	.	·00002857
Tin	.	.	.	.	·000021730
Silver	.	.	.	.	·000019097
Zinc	.	.	.	.	·000029417

The coefficients for cubical expansion are three times those for linear.

To find the weight of a gas in lbs. per cubic foot at standard temperature and pressure, multiply half the molecular weight by ·00557.

## II. MISCELLANEOUS QUESTIONS.

["B. of E." denotes that the question has been given at the Board of Education Examination in Mining. Questions marked "C. M." have been set at various examinations for Colliery Manager's certificates.]

1. What gases are met with in mines? What are their chief properties?
2. Explain the action of the siphon.
3. How does the gauze cylinder in a safety lamp act as a preventive of explosion?
4. What is meant by gaseous diffusion?
5. What is black-band ironstone? How does it differ from hematite iron ore?
6. Explain how a siphon is sometimes used for draining a mine, and state conditions which limit its usefulness.
7. Describe the Clowes Hydrogen lamp, and state the amounts of fire-damp which can be recognised by it.
8. Three specimens of fuel A, B, and C gave when analysed the results shown in the table below. What would you call each of these fuels? Note any peculiarities of composition which strike you as likely to affect the calorific value of the fuels. (B. of E.)

	A	B	C
Carbon . . . .	36·0	44·8	39·58
Hydrogen . . . .	2·55	3·9	3·19
Oxygen . . . .	12·16	10·6	16·35
Nitrogen . . . .	·57	·6	·45
Sulphur . . . .	—	·7	—
Water . . . .	30·79	33·9	32·26
Ash . . . .	17·93	5·5	8·17

9. State what you know concerning the manufacture of briquettes or patent fuel.

10. Which will diffuse the more quickly, black-damp with air, or fire-damp with air?

11. A mixture of fire-damp and air measures 20,000 cubic feet and contains 8 per cent. of fire-damp. How much air will have to be added so as to form a mixture which will not show a cap with an ordinary safety lamp?

12. A bunker for storing coal measures 20 feet long, 15 feet wide, and 10 feet deep. How much coal will this contain of specific gravity if the latter be taken as 1.3?

13. What reasons are there for concluding that coal dust is a most dangerous element in coal mining? State when it is most and least dangerous.

14. What is the composition of anthracite, cannel and bituminous coals, and of coke?

15. Describe the principle of the safety lamp.

16. A piece of coal weighing 20.65 ounces when immersed in water weighs 4.13 ounces. What is its specific gravity and its weight per cubic foot?

17. Why is water such a useful medium for heating purposes?

18. What would be the effect upon the height of the mercury column of an ordinary barometer if taken down into a deep mine?

19. How can air be proved to have weight? How would you compare the relative weights of various gases?

20. Describe the construction and use of an ordinary mercury barometer. If water or glycerine be used instead of mercury, how is the length of the column effected? What is the advantage in using glycerine?

21. The tube of a thermometer has a very narrow bore, but the bulb has a large one. Why is this?

22. Why is mercury used in making thermometers? Is any other liquid used for this purpose? If so, why?

23. How can it be proved that solid liquids and gases expand when heated?

24. Explain how the boiling point and freezing point are determined on an ordinary thermometer, and how it is graduated.

25. Explain the terms "Conduction," "Convection," "Radiation" of heat, and "Specific Heat."

26. The specific heat of mercury is said to be .033. What does this mean?

27. Describe an experiment to prove that water is a bad conductor of heat.

28. 1 lb. of water and 1 lb. of iron are heated to the same temperature. They are then both placed on a block of ice. Which will melt most of the ice, and why?

29. How much fire-damp would have to be given off to foul a current of 6000 feet of fresh air per minute sufficiently to allow the presence of the gas to be detected by the flame of a safety lamp? (C. M.)

30. State the most important laws and principles affecting the ventilation of mines. (C. M.)

31. Account for rise of the air in an upcast shaft, and show the relation of the movement to the action of the siphon. (C. M.)

32. Describe and sketch a safety lamp suitable for a fiery mine, and able to withstand a velocity of 30 feet per second. (C. M.)

33. Describe, with sketches, the form of safety lamp which you consider best for use in fiery mines with a high velocity of ventilating current, also lamps with which you are acquainted for testing for small quantities of gas. (C. M.)

34. What explosives would you employ in a fiery mine? Give details of when and how you would use it. (C. M.)

35. What, in your opinion, is the best explosive now in use in coal mines? Give your reasons, stating its leading characteristics; mention any disadvantages, and say what are your views as to substitutes for explosives. (C. M.)

36. State what you know about the explosive properties of coal dust, and what precautions you would take in working a dry and dusty mine. (C. M.)

37. When fire-damp at its most explosive point is fired,

what takes place? What are the resultant gases, in what proportions, and what are their properties? (C. M.)

38. Describe a method, or methods, of testing for small percentages of fire-damp, such that the presence of gas cannot be detected by the means of an ordinary safety lamp.

39. Name the chief constituents of pure air, and give their chemical properties. Give the chemical properties of fire-damp and choke-damp, and their weights compared with atmospheric air.

40. What effects may coal dust produce in a mine, and how would you guard against them?

41. What gases are given off in a coal mine from spontaneous combustion or gob-fires, from the coal itself, from lamps, and from shot-firing? What are their properties and chemical compositions?

42. What do you understand by the "pitting" of boiler plates, and how is it caused?

43. Which is the most likely part of a working place to find fire-damp accumulating? What is the most explosive mixture of fire-damp and air? What effect has fire-damp, unmixed with air, upon the flame in a lamp?

44. How would you detect the presence of fire-damp, and how would you estimate the proportion of gas present? Describe how you would search for this gas in a drawing-road and in a working place. (B. of E.)

45. What is the effect of coal dust on the atmosphere of fiery mines, and what special precautions are required in working such mines? (B. of E.)

46. Describe the Clowes lamp, and state how small a percentage of fire-damp can be detected with it.

47. Describe the general principles of "pneumatophores," and the conditions which have to be fulfilled to make the appliances of practical use.

48. What is the cause of acid water in mines, and what precautions are necessary for raising it? (B. of E.)

49. Given the analyses of a number of samples of coal-carbon, hydrogen, oxygen, and nitrogen, calculated as per-

centages, independently of ash, sulphur, phosphorus and moisture, and the moisture given as a percentage of the whole, how would you be able to distinguish in a general way between lignite, long-flaming bituminous coal, caking coal, semi-bituminous coal, and anthracite? (B. of E.)

50. Express the chemical composition of the following gases by means of the usual symbols: Carbon dioxide (carbonic acid gas), carbon monoxide (carbonic oxide), marsh gas (fire-damp), sulphuretted hydrogen. Under what circumstances may one or more of these gases occur in a mine? How can the presence of each be detected in a summary manner? What effect is produced if any one of them is allowed to accumulate in a mine in which naked lights are used? (B. of E.)

51. What is the composition of (a) dynamite, (b) blasting-powder? What is the essential difference between their modes of ignition? (B. of E.)

52. What are the two most deleterious gases most commonly met with in coal mining; their chief properties, special danger from each; in what places are they most likely to be found, and how is their presence detected? (C. M.)

53. What is atmospheric air? Describe its chemical composition; give weight of a cubic foot of air when temperature is 32°. Describe uses of barometer and thermometer. (C. M.)

54. What is the specific gravity of nitrogen gas? Describe its properties and state its effect on the oxygen of the atmosphere. (C. M.)

55. What is oxygen gas? State specific gravity and various properties of the gas. (C. M.)

56. What is carbonic acid gas? Give the different names of this gas in connection with mining, and state how it is produced. State the chemical composition of this gas, viz., by atoms and by weight. What percentage of this gas mixed with air renders it dangerous to life to breathe, and to extinguish lights? What is the specific gravity of the gas, and where is it likely to be found in mines and shafts? (C. M.)

57. Proto-carburetted Hydrogen.—What is this gas called in mining; what is its chemical composition, by atoms and by weight, and its specific gravity? Can it be breathed in the pure state; what percentage of this gas is required to be mixed with air to enable you to detect its presence in the flame, and what proportion of air to the gas renders the explosive force greatest? State what gases, when present, lessen the explosive force of the gas. Where would you expect to find this gas in a mine? (C. M.)

58. What is after-damp? What are the gases contained in this gas? Will this gas support either life or combustion? (C. M.)

59. What is carbonic oxide? State the composition by atoms and weight. What is the specific gravity; what effect has it on animal life; is this an inflammable gas, and does it support combustion; how is this gas produced? (C. M.)

60. What is spontaneous combustion, or gob-fires in a mine; how are they produced? (C. M.)

61. Suppose in making an examination of workings, wastes or old roads, you discovered a gas near the roof of a mine, which extinguished your lamp without exploding, what gas would this be? (C. M.)

62. Explain the action of the siphon, and its use and application in draining mines. (C. M.)

63. Describe the gases you expect to meet in a mine. (C. M.)

64. What part of a mine would you expect to find carbonic acid or black-damp, and how would you deal with a body of the gas? State its effects on animal life. (C. M.)

65. What is carbonic oxide? How is this gas produced in mines, and what effect has it on animal life? (C. M.)

66. State fully how you would examine a place for gas. What is the smallest proportion that will show a cap? At what proportion does it begin to explode, and what proportion will extinguish a light? (C. M.)

67. If the air in a mine contains enough fire-damp to show a "cap" which is only just perceptible, what additional per-

centage of gas will suffice to make the mixture highly explosive? (C. M.)

68. How do spontaneous fires occur in mines; how is their presence first detected? (C. M.)

69. State your experience with the barometer and thermometer in connection with mining. Suppose you have a mine 100 yards in depth and another 250 yards in depth; what will the difference in the reading of the two depths be? What is a thermometer, and how is the ventilation of the mine affected by the changes of this instrument? (C. M.)

70. What is atmospheric air? Describe its chemical composition. (C. M.) Is there any variation of this air? If so, how does it affect the ventilation of mines?

71. State your experience in regard to blasting in fiery and dusty mines, and in non-dusty mines. Name the explosive you have been accustomed to. What examinations are necessary before blasting in gassy and dusty mines? (C. M.)

72. What is the difference between furnace and fan ventilation? (C. M.)

73. What is a "dynamo" and what is a "motor"? (C. M.)

74. In what way may danger arise from coal dust in collieries? (C. M.)

75. If 100 cubic inches of air weigh 31 grains Troy under a barometric pressure of 30", what will 100 cubic inches weigh at a higher altitude, where the barometer stands at 26" without any change of temperature? (C. M.)

76. To what extent will 2000 volumes of gas expand if the barometric pressure fall from 30.15" to 29.30"? (C. M.)

77. Certain mixtures of fire-damp and air explode feebly, while others again are highly explosive. How is this so? Give chemical formula for a perfect explosion. (C. M.)

78. A volume of air measures 30,000 cubic feet, and consists of 79 per cent. nitrogen, 20.9 per cent. oxygen, and .1 per cent. carbon dioxide. How many cubic feet of each gas is present?

79. Enumerate the noxious gases found in collieries, and state the composition of each. (C. M.)

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